STANDARD CHEMISTRY

[illegible]

100 A

TABLE OF EQUIVALENTS

1 inch = 2.54 centimeters.
1 foot = .3048 of a meter.
1 mile = 1.6093 kilometers.
1 centimeter = .3937 of an inch.
1 meter = 1.0936 yards.
1 cubic inch = 16.387 cubic centimeters.
1 cubic foot = 28.317 cubic decimeters.
1 cubic centimeter = .061 of a cubic inch.
1 liter = 1.0567 liquid quarts.
1 liter = .908 of a dry quart.
1 square inch = 6.452 square centimeters.
1 square foot = .0929 of a square meter.
1 square centimeter = .155 of a square inch.
1 square meter = 1.196 square yards.
1 grain, Troy = .0648 of a gram.
1 ounce, Avoirdupois = 28.35 grams.
1 ounce, Troy = 31.104 grams.
1 gram = .03527 of an ounce, Avoirdupois.
1 gram = .03215 of an ounce, Troy.
1 gram = 15.432 grains, Troy.
1 kilogram = 2.2046 pounds, Avoirdupois.

TABLES OF WEIGHTS AND MEASURES

TROY WEIGHT

24 grains (gr.) = 1 pennyweight . . pwt.
20 pennyweights = 1 ounce . . . . oz.
12 ounces = 1 pound . . . . lb.
1 pound = 5760 grains.
1 ounce = 480 grains.

AVOIRDUPOIS WEIGHT

16 ounces = 1 pound . . . . lb.
100 pounds = 1 hundred-weight . . . . . . cwt.
20 hundred-weight, or 2000 pounds = 1 ton . . . . T.

APOTHECARIES' WEIGHT

20 grains = 1 scruple . . . . . . sc.
3 scruples = 1 dram . . . . dr.
8 drams = 1 ounce . . . . oz.
12 ounces = 1 pound . . . . lb.

LINEAR MEASURE

12 inches (in.) = 1 foot . . . . ft.
3 feet = 1 yard . . . . yd.
5½ yards, or 16½ feet = 1 rod . . . . rd.
320 rods = 1 statute mile . . . . mi.
1 mile = 320 rods, or 5280 feet, or 633,360 inches.

METRIC LINEAR TABLE

10 mil′li-me′ters (mm.) = 1 cen′ti-me′ter . . . . cm. = $\frac{1}{100}$ M.
10 i-me′ter . . . . dm. = $\frac{1}{10}$ M.
10 er . . . . M.
10 a-me′ter . . . . Dm. = 10 M.
10 to-me′ter . . . . Hm. = 100 M.
10 -me′ter . . . . Km. = 1000 M.
10 ′ia-me′ter . . . . Mm. = 10000 M.

LE OF WEIGHTS

cg. = $\frac{1}{100}$ gm.
dg. = $\frac{1}{10}$ gm.
gm.
kg.
measure of capacity.
L.

Med
K921

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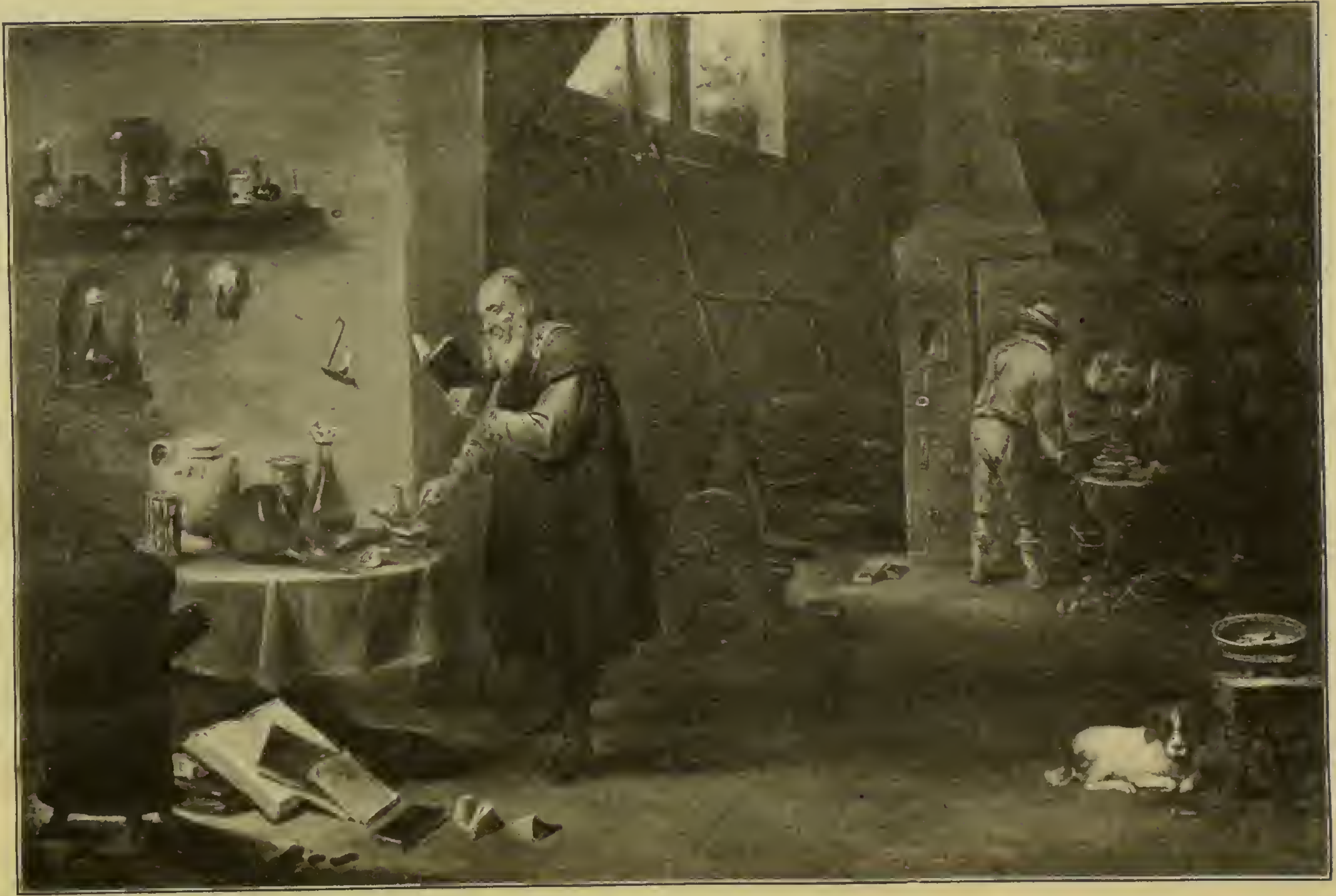

AN ALCHEMIST
by
TENIERS THE YOUNGER

A MEDIÆVAL STUDENT OF CHEMISTRY

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ELEMENTARY CHEMISTRY

LIBRARY OF THE BRITISH MEDICAL ASSOCIATION

By

HOLLIS GODFREY

HEAD OF THE DEPARTMENT OF SCIENCE
GIRLS' HIGH SCHOOL OF PRACTICAL ARTS
BOSTON, MASS.

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To
MY MOTHER

PREFACE

Four ideals have governed the writing of this book. The author has desired to obtain simplicity; to reach the understanding of the student; to rouse the pupil to a realization that the science of daily life is identical with the science of the school room; to include all the essential facts and theories which could be rightly assimilated in one year's work in elementary chemistry.

To obtain logical results with such intentions, a proper sequence of ideas is of prime necessity, and the pupil, once falling into the general trend of the work, should have a clear uninterrupted progression of thought from the beginning to the end of each chapter. No book which is a mere encyclopedia of facts arranged without reference to their teaching value can produce a maximum of effect. Indeed this volume makes no attempt to be an encyclopedia of chemistry. Every chapter represents far more matter rejected than it holds of matter accepted. It has been a constant purpose to bring forward wide-reaching general truths in the form in which they would most effectively impress the student. Technicalities, statements which require hundreds of words of explanation to a few words of definition, have been omitted and in every possible case illustrations and examples have been chosen from those things with which the student is likely to be most familiar. The writer, recognizing the difficulty of his task, welcomes criticism, suggestions, or inquiries concerning the adaptability of this book to individual needs. Mention should also be made here of the fact that the data obtained do not always agree with results found in many standard texts. For example, the figures on the percentage of nutrients present in milk, which come from a widely extended recent research on various grades of milk have seemed more satisfactory than the quantities generally given. Where modern research has seemed to differ from previously accepted statements, the claims of each side have been carefully considered and deliberate choice made.

This book is divided into three parts. Part I consists of thirty-four chapters, and treats of those portions of elementary inorganic and organic chemistry which the author has considered of most value to the student. Each chapter, with the exception of the final chapter on chemical arithmetic, may represent the work of a single week. The chemical arithmetic has been grouped in one place in the belief that this branch of the subject is of more value when treated as a whole than when arbitrarily inserted in scattered fragments. Part II is made up of brief epitomes of the thirty-three descriptive chapters of Part I. Epitomes have often proved their value as bases for review. Part III consists of questions drawn from the individual chapters. The pupil studying a lesson at home is far more likely to obtain a clear conception of his work if guiding questions are at hand. The text as a whole can be used with any laboratory manual. A laboratory manual, however, has been prepared to accompany the work. It endeavors to show the same theory of the connection of natural science with the living world which the book is designed to bring forward.

To the readers of this book the author owes a deep and sincere gratitude for many kindly and helpful suggestions. The whole manuscript has been read by Prof. James F. Norris, head of the School of Science of Simmons College, by Miss Laura White, head of the Department of Science in the Girls' High School of Boston, by Frederic W. Howe, head of the Department of Chemistry in the Framingham Normal School, and by Miss Josephine Hammond, head of the Department of English in the High School of Practical Arts in Boston. The manuscript has been read in part by George D. Bussey, head of the Department of Science in the East Boston High School, and by F. M. Eaton, Consulting Chemist of Oakland, California. The more purely theoretical chapters have been read by Dr. Arthur Blanchard, Assistant Professor of Chemistry at the Massachusetts Institute of Technology, and the proof has been read by George Owen, my former assistant, now of Harvard College, by Newman B. Gregory, of the Massachusetts Institute of Technology, and by Miss Charlotte Coffin, my present assistant.

The illustrations in this book have been gathered from numerous sources by the kindly efforts of a number of people. The author is specially indebted to Baldwin Coolidge, of Boston, to Frederic W. Howe, to the Walker Gordon Laboratory Co., and to the Brooks estate. The figures on pages 75, 102, 125, 145, 155, 191, 199, 202, 222, 231, 233, 245, 320, 321, 322, 334, 354, are reproduced from Newth's Text-Book of Inorganic Chemistry by permission of Mr. Newth; that on page 46 from Frankland and Japp's Chemistry by permission of Messrs. J. and A. Churchill; while for those on pages 214, 216, and 217, the author is indebted to the courtesy of *The Paper Mill and Wood Pulp News.*

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ELEMENTARY CHEMISTRY

THE REASONS FOR THE STUDY OF CHEMISTRY

The eye, that camera which gives us our chief knowledge of the world, sends its messages to our brains by a series of tiny, sensitive nerve beginnings called the rods and cones which form the inner lining of the retina, the eye-screen on which wonderful pictures are thrown. All over the retina, from top to bottom, from side to side, these small receivers are constantly flashing us all the world of beauty about us, sunset and sunrise, sea and mountain, great picture and noble building. In one spot only are they absent, the blind spot of the eye. Every man, woman, and child in the world possesses that one blind spot of the eye. No rod or cone exists at this spot to receive image and form. Our sight depends wholly on the action of the rest of the retina. Though we can see the whole field of vision with both our eyes, yet each alone finds one place hidden from its view because of that one sightless space.

The blind spots of our eyes are nothing, however, to the blind spots of our mental vision. To the ordinary passerby, a bird-filled copse may seem alive with common sparrows, when to the bird lover there may appear a bewildering train of warblers whose beauty and whose voices give constant pleasure. To the musician the harmony of a great symphony gives a joy unknown to the man whose soul is not attuned to music. The great bridge to whose construction the crossing thousands give not a thought shows to the engineer weeks and months of never-wearying toil, thousands of computations and exact measurements. The bird lover, the musician, and the engineer have each cleared off some blind spots from their mental vision by long study, careful thought, and trained observation.

The province of chemistry is to clear the blind spots from our mental vision as regards the meaning of the immediate world in which we live. This is the science which enables us to understand the air which we breathe, the water which we drink, the soil on which we stand, the food which we eat, and the clothes which we wear, the carriages in which we ride, and the houses in which we live. In a word, chemistry is the science which explains the every-day things of life. Its study is a necessity since no man or woman can live rightly unless he or she knows something about the uses and composition of the necessities and luxuries used in our daily life in this beginning of the twentieth century.

The great majority of us owe our livelihood indirectly at least to this science. Take the chemistry of a single element, iron. It would be hard indeed to gather together a company one fourth of whom did not depend for daily food and shelter on the continued production of iron, — a production made possible by the chemist. He is the man who tells the value of iron ore, how iron can be separated, how it can be prepared, and how it can be used. Think for a moment of the myriad occupations which depend on iron. The railroads of the country give employment directly and indirectly to workers whose numbers reach to hundreds of thousands. The railroads depend on iron for rails, for wheels, for the construction of the cars. The machine shops of the cities are sending out millions of machines whose operation depends on iron. The men who make the machines and the men who use them, the men who run the factory and the men who use the product of the machine, whether that product is shoes or clothes, furniture or books, all depend on the present outpouring of iron for their livelihood. Is not some study of iron, some clearing of the mental vision in this respect wise?

As the carpenter needs to know the different uses of the rip-saw and cross-cut saw, in order that he may saw apart his boards, as the blacksmith needs to know the use of the bellows which blow his fire till its glowing heat will bend the horseshoe, so we need to know something of the construction and the use of those material substances on which our livelihood depends. And knowledge of the properties of one substance will not suffice. This age of ours is a very complex one. A thousand different substances are called daily from the ends of the earth to give us necessities or furnish us luxuries. They weave back and forth in the shuttle of our daily life, and the more we know of them, the more the world beautiful about us opens to our view. To gain such increased knowledge we must rely largely upon the science of chemistry.

Take a second example, the work of the chemist in adding thousands of by-products to the wealth of the nation. The very word by-product was scarce known a few decades ago. To-day it appears far and wide in the daily press. Strictly defined, "a by-product is something produced, as in the course of manufacture, in addition to the principal product." But, in case after case to-day, the by-product, not the principal product, makes manufacture profitable. In soap-making the competition has grown so great that the old methods of bringing together fats and lye to make soap, keeping the soap, and throwing everything else away would mean great loss. The saving and use of the glycerine which is produced with soap makes a profit which in some cases is greater than that made from the soap itself. The making of baking soda by the old process is continued largely for the sake of the by-product, muriatic acid, once thrown away, now saved and sold for use in a thousand other manufactures. More and more the chemist is recognized as the man who can save the waste that was once lost, who can find means to utilize the most unpromising refuse and bring profit from some heretofore neglected by-product.

Chemistry is doing many things: by its study of the needs of the soil it is making two blades of wheat grow where one grew before; by its study of the metals it is furnishing new industrial worlds with means of transportation, of manufacture and communication; by its study of the earth it is giving the inhabitants of the world great stores of the products which come from plant and tree, from coal and mineral; thereby producing such wealth that the poorest mechanic to-day lives a far more luxurious life than did the kings of a few hundred years ago; last but not least, its applications to preventive medicine are safeguarding thousands of communities from disease. Here opens the study of this great science.

CHAPTER I

THE MAKE-UP OF THE WORLD

Methods of studying chemistry. — The cabinet-maker, before he can produce carved cabinets of rare woods, the diamond-cutter, before he can cut the facets of precious stones, the bookbinder, before he can turn out bindings which will delight the eye, must learn to know his materials and his tools. He must go through some toilsome labor before he can do artistic work. Before we can explore the wonders that chemistry spreads before us, and study gases and gleaming metals, plant stalk and white beach sand, we must know something of the way all these things are made; we must learn what is the basis

Earth Compounds

This picture, like the two following, shows many types of earth compounds. Sand, rock and clay, water and upper soil, metallic and non-metallic substances, appear.

of the structure of these substances. How do they change and vary? What methods may be applied to gather, to test, and to examine them? The next two chapters are devoted to a consideration of these things.

The visible and invisible in the universe. — Look up into the heavens on a clear August night, when the sky is studded with myriad points of light, and try to count the stars with the naked eye. You may count on hour after hour, scarcely registering a fraction of their number. Yet should you look through a common telescope, thousands on thousands of new stars would spring into your field of vision and still more powerful telescopes would reveal other hundreds of thousands of stars unseen before. Since no limit to these heavenly lamps appears it is at least a natural belief that far beyond the range of the largest telescope millions of uncounted stars may shine of whose existence we may never definitely know.

Units of the universe and of matter. — The universe is so infinitely large that we can see some portion of the units or parts which compose it in the heavenly bodies, that is, see many of the individual suns and some of the planets which belong to them. It is only because the universe is infinitely large that its units, which we see, appear to us as great. In comparison with the whole they are very small. The matter which makes up the different units of the universe — the gases of Saturn, the dead surface of the moon, or the tree-clad hills of the earth — are bodies which are infinitesimally small compared with the suns or planets of which they are a part. This matter is furthermore made up of units or parts so infinitely small that we cannot see them with the most powerful microscope any more than we can see the outer boundaries of the stars with the most powerful telescope. Since we cannot see the units or particles which together make up the matter of water, or soil, or wood, or leather, but only the masses which result from the putting together of such units, we cannot tell exactly what these units are, nor can we state as an absolute fact the way that matter is put together, but rather as a hypothesis.

Two words in that sentence, "matter" and "hypothesis," need some explanation.

Matter. The definition of a hypothesis. Matter may be defined as something that can be felt, can be moved, can be weighed. Whole books have been written in the attempt to define that word exactly, but we all know practically what it means, and that those things which we can see or touch are matter. The wind shaking the tree, the leaf rustling in the breeze, or the rain pouring on the roof, all these are matter, all these can be felt, moved, or weighed. Now as to fact or law as opposed to hypothesis. Most of the changes which are considered in chemistry are facts. We can see them with our eyes, test them with our senses, prove that they are correct because they act in the same way a thousand times running. From some of these facts we can make laws. To illustrate: We know that gases expand when heated, and contract when chilled. Over and over again this has been proved and never in the history of man have gases done anything else. In consequence we may state a law which governs the expansion of gases. Every more powerful telescope, on the other hand, which has been built, has given us thousands of new stars more than the one which came before. We can make a hypothesis based on these results saying that there are many stars beyond our vision. There is every reason to believe that this is so, but we cannot prove it. We cannot see the distant stars, but all the evidence in our hands goes to prove that they are there. Therefore we can form a hypothesis that they exist; and a hypothesis is the statement of a belief. It is equally possible to form a hypothesis concerning the composition of matter.

The atomic hypothesis. — The hypothesis or belief which many scientists hold with regard to the composition of matter considers that it is made up of an infinite number of tiny systems similar in many respects to the great planetary systems of the universe, similar, for example, to our own planetary system where the Earth, Mars, Saturn, and the rest revolve about the sun. A bolt of iron, dead, motionless, and entire as it seems, is, nevertheless, according to this belief, made up of billions upon billions of units which are in vibration, and which, in turn, are made up of yet smaller particles. The units, which like bricks build up the mass of the iron, are called molecules. The smaller particles which build up the molecules are called atoms. As our earth is one of the planets of our solar system, so the atom may be considered one of the planets of a molecular system; and as our whole solar system, the sun, and all the planets, combines with the total millions of other solar systems to make up the sum of the universe, so the sum of the molecules, made up of the atoms, makes up the total mass of any body, be it wood or iron, water or quicksilver. The mass is the whole, made up of molecules; the molecules are the units, made up of still smaller parts, the atoms. It is impossible to see even the largest of the units, the molecule. The smallest grain that the most powerful microscope discloses contains millions of these tiny particles. To use the brilliant example of one scientist, "If you took a single drop of water and could enlarge it to the size of the whole earth, each of the molecules, of the tiny units, that make it up would then be about the size of a baseball." Think of that proportion. The molecule is to the drop of water as a baseball is to the whole earth. However firmly we may believe in its truth, we are not very likely to prove this hypothesis, which is known as the atomic hypothesis, and so change it from hypothesis to law.

Indivisibility of the atom. — When the atomic hypothesis was brought forward to account for the composition of matter, it was believed that the atom was indivisible. To-day that is not generally believed, but we do know that in almost all cases in chemistry the atom acts as if it were indivisible, and in no ordinary chemical action does it separate. In consequence we may talk about the atom exactly as if it could not be divided. This indivisibility of the atom produces some rather striking results. If an atom cannot be divided, it comes about as a result that every atom of the same substance is alike wherever it is found. The iron atom, for example, is the same in the wall of a great battle-ship as in the red clay, which may be behind our house.

Elements. — There are only between seventy and eighty different atoms, so far as we know, and when these atoms combine, like with like, to form molecules, there are only between seventy and eighty different elementary molecules, (excluding certain forms known as allotropes of which we shall have occasion to speak later) molecules all of whose atoms are the same. Consequently, since masses are made by the union of molecules, there exist only between seventy and eighty perfectly simple substances which cannot be changed into any two different parts, no matter how hard you try to separate them. Such simple substances are called elements. All atoms are elementary because they are of necessity simple.

Compounds. — To form an elementary molecule a group of sulphur atoms may be gathered together. They would make up a sulphur molecule, which, joined to other sulphur molecules, might build up a bar of roll sulphur. No matter how completely you separated that bar of sulphur, you could never get anything else than sulphur from it. It is wholly built from elementary units. If on the other hand some sulphur atoms joined with some oxygen atoms to form a compound known as sulphur dioxide, then you would have a compound molecule which, with other like molecules, would build a compound mass. If we treat such a compound properly it will separate into its elementary parts. In this case the sulphur dioxide will go back to the sulphur and oxygen with which it began. To sum it up, we may say that an element is a substance simple in itself and which cannot be divided into anything simpler. A compound is a combination of elements which can always be divided into parts simpler or more elementary than itself.

The number of elements. — A large number of the elements play little part in the world's work. Most of the substances that we use in a long life are made up from less than thirty of them. The full list may be found on the cover page, but for the present it is enough to say that iron and copper, and zinc, gold, silver, mercury, carbon, phosphorus, and sulphur, hydrogen, oxygen, and nitrogen, are elements. In this little list of twelve may be found the constituent parts of many of the compounds of our whole globe. A full list containing the names of all compounds known to-day would fill many books. It is enough to speak of water and iron rust, of clay and of sand. All the molecules of these substances are built up by the additions of different atoms. The matter which is built up from these tiny particles may appear in masses of any size, raise the loftiest mountain or form the whole crust of this earth. The wide-spreading ocean that covers more than

Earth Compounds

half of the globe is made up almost wholly of the molecules of one compound, water, while the atmosphere which spreads around and above us is chiefly a mixture of the molecules of two elementary gases, oxygen and nitrogen.

Three states of matter. — In those three examples, the earth's crust, a solid, the sea, a liquid, and the atmosphere, a gas, we have the three states of matter. A vast number of known substances exist in all three states. They may exist as solid, liquid, or gas. The form which is taken depends on the two factors of temperature and pressure. Water when exposed to winter cold appears as solid ice, at ordinary temperatures is liquid, and when put on the stove boils away in the form of gaseous water or steam. It is water in every form. Its molecule, the unit which makes it what it is, remains wholly unchanged: cold condenses steam to liquid water, and a more intense cold produces ice: heat on the other hand brings the reverse changes, melting the ice to liquid water, and then producing steam.

Effect of temperature and pressure upon matter. — Solid, liquid, and gas, the three states of matter, run through the whole world of nature. Though every substance does not appear in every form, so many of them do that we have no reason to believe that all of them might not, were the right conditions once obtained. Pressure and temperature determine generally in which of the three states the body appears, and one of the first questions that we ask about any new substance, be it element or compound, is with regard to the state in which it appears at ordinary temperatures and pressures. Of the elements, oxygen, nitrogen, hydrogen are three typical gases. Sulphur, silver, and phosphorus are three typical solids, and mercury (quicksilver) a typical liquid.

Metals and non-metals. — Though the elements can be classified as regards the state of matter in which they appear at ordinary temperatures, it is better to classify them along a somewhat different line, into metals and non-metals. We know roughly that the copper of the telephone wires, the iron of railroad tracks, and the silver of a coin or a spoon all belong to the class, metals, and that the general characteristics of such substances as sulphur or charcoal (which is really the element carbon), which are non-metals, are, as a class, quite opposed to those of the metals. Suppose we consider some of the points wherein these groups differ.

Difference between metals and non-metals. Metalloids. — Take the metals first. If you put a silver spoon into hot water, the heat soon runs up the handle. That metal is a good conductor of heat. The warmth travels rapidly from one end of the spoon to the other. A stick of charcoal, on the other hand, might be at boiling heat on one end, in a glass of boiling water, and you would feel no warmth at all at the other end. That contrast holds throughout. The metals, in general, are good conductors of heat; they pass heat along readily from one part of their substance to another. The

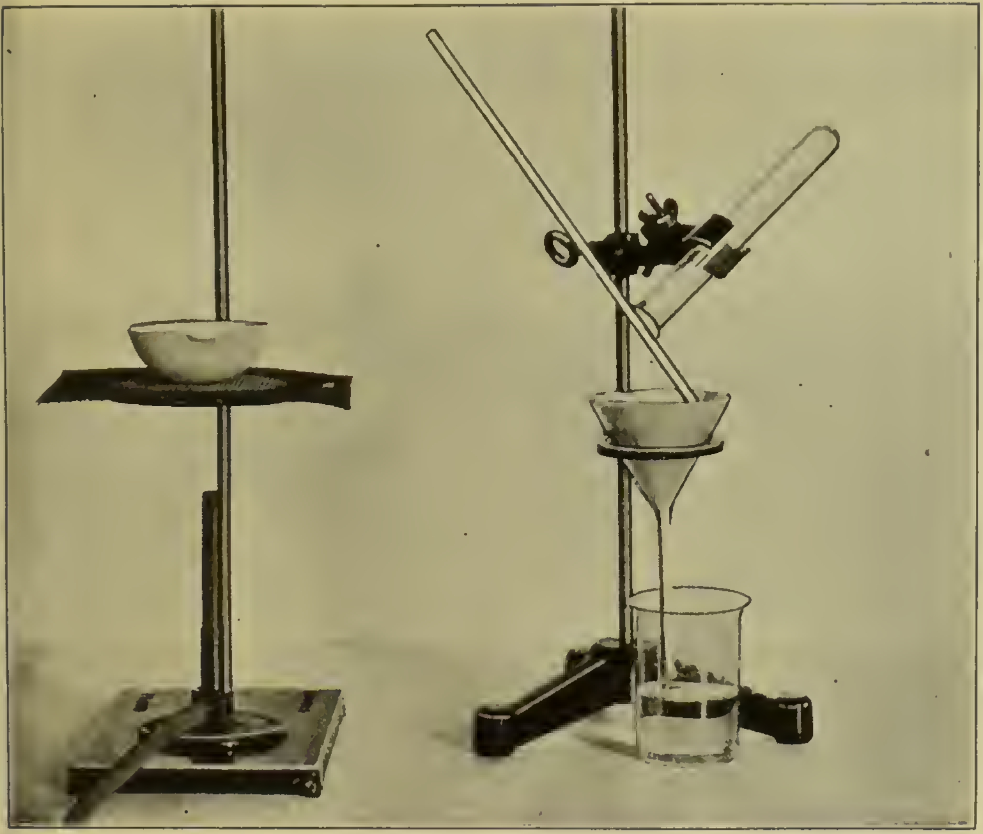
Earth Compounds

non-metals conduct heat only with difficulty. The telephone poles beside the road bear wires of copper over which flash messages across thousands of miles. Copper, like most other metals, is a good conductor of electricity. Metals possess a peculiar shiny luster, which is so distinctive in appearance that we have an adjective describing it, the word metallic. To keep the electric flow from escaping from carrying wires glass insulators are used. Glass is a non-conductor of electricity, and like glass most of the non-metals are either non-conductors or poor conductors of the electric current. Non-metals have generally a dull or earthy appearance in the solid form. These are the chief external differences of the two classes: for the clearest definition of the difference between them, we must wait until our study of the acids. Metals and non-metals are by no means divided from each other by any hard and rigid line. There is a group called the metalloids which stand on the boundary between the two, and sometimes act like metals, sometimes like non-metals. Despite the trouble we sometimes have in classifying these classes exactly, this classification into metals, non-metals, and metalloids serves many useful purposes and we shall often be obliged to refer to one or another of the groups.

Physical and chemical change. — A French philosopher is credited with saying in new form what the Greek Heraclitus said 2500 years ago, — "Man's life is spent in change. His work and play consist of changing something." He was not far wrong. When we throw a ball, or cross the continent on a railroad train, we have one type of change, when we burn coal and have left hot gases and gray ashes, or take sweet cider and ferment it to sour vinegar, we have the other. In all the cases we are changing something, but in these illustrations we find two kinds of change wholly distinct, the first physical, and the second chemical. The ball that was thrown was changed in no way by its motion through the air. We arrive in San Francisco substantially the same person as when we left New York. The change is physical. There has been no differentiation of properties. Our body has merely altered its position: it has not in any way become a new thing. Compare such a change for a moment with the burning coal or the turning cider. No jugglery will bring the coal back from the parts into which it has turned, the ashes and the gases. The sour vinegar has little relation to the sweet cider from which it comes. These are chemical processes. New substances have formed by these changes which are radically and totally different in their properties from the substances which produced them.

Differences between physical and chemical changes. — Here is the essential difference between physical and chemical change. Physical change may not only transfer position, it may totally alter appearance: steam, for example, is wholly unlike the ice from which it may come, but whatever guise a body may assume under a physical change its real chemical composition is not altered. Sugar may disappear dissolved in water, but let the water once be evaporated and the sugar reappears, white, sweet, and granular as before. Solution of this type is purely a physical change. No substance formed by a chemical change can ever be brought back to its original condition by a physical change. The cider changes to vinegar; black iron left on the roof turns to a streak of red iron rust; chemical changes such as these turn the substances acted upon to other substances, wholly different in their characteristics and different in their action. Drop a bit of metal into an acid and it dissolves as did the sugar, but unlike the sugar it never reappears in the same form. Instead of the bright metal, a powder of wholly different characteristics, a compound of the metal and the acid, appears when the acid is evaporated. A chemical change has taken place and a new substance has been formed.

Physical changes then involve no real change in structure. Chemical changes involve radical changes in structure.



Evaporation and Filtration

Two common processes used in the experimental consideration of physical and chemical changes are pictured here. On the left a solution is being evaporated (boiled away, we should say in the kitchen), that is, the liquid of the solution is changing to a gas under the influence of the heat produced by the lighted gas coming from the gas burner below. (This type of burner is called a Bunsen burner.) To keep the evaporating dish from cracking it is placed upon an asbestos centered wire gauze during the operation. On the right a liquid holding a solid (like water holding sand) is being poured from a test tube, down a glass rod and into a glass funnel lined with filter paper. Filter paper is porous and allows liquids to pass through while it holds solids back. The liquid runs down into a beaker below. Such a separation is called filtration. The solid which remains is called a residue, sometimes a precipitate (we shall explain that word later). The liquid which passes through is called a filtrate.

CHAPTER II

OUR KNOWLEDGE OF THE MATTER OF THE WORLD

The five senses in chemistry. — Our five senses, sight, touch, smell, hearing, and taste, are the only messengers to our brain from the world without. On them we are obliged to depend for all the information which we possess. Indeed the sum total of our knowledge comes from the varied sensations which the senses transmit to our brain, so it is inevitable that in our study of chemistry we should be guided in every test by the action of these brain messengers, and should give the properties of each element and compound in terms of their every-day use.

Testing common salt. — Suppose, for instance, we desire to study common salt. Our eyes tell us that we have before us a white powder, which dissolves when we throw it into water, which colors a flame yellow when thrown on a burning substance. We can see also that a white solid falls when a silver compound in solution is added to a solution of salt. Our tongue tells us of the peculiar salt taste. The nose tells us that the substance is odorless. The hearing gives us no clue, but the sense of touch is replaced by that invaluable aid to the chemist, the balance; weighing it we can learn whether it is lighter or heavier than other substances. In considering the properties of any substance we cannot be satisfied that we know them all until we have exhausted the possibilities of the five senses in relation to the substance.

Changes signified by differences in weight. Analysis, synthesis, and metathesis. — On the comparison of weights, the amount that a substance gains or loses during a chemical action, constantly depends the answer to the question, "Has the body changed from one substance to another?" In almost every case where a body changes in weight, such a change in weight indicates a chemical as opposed to a physical change, excepting, and these are most important exceptions, the cases where a body changes its state of matter, and, turning into the gaseous state, passes off into the air, or when a body dissolves, or absorbs some liquid or gaseous matter from the atmosphere or from the receptacle which contains it. A loss in weight commonly means that a body has separated into simpler parts. A gain in weight commonly means that a body has added something to itself. In consequence of these general principles we constantly use the balance to tell us whether or not a chemical change has taken place, and, if it has, to tell us also, just how much of a change has occurred, exactly the weight added or taken away. The balance is one means by which we can generally prove that these changes have occurred. When a compound is separated into its component parts, the process is called analysis, a word signifying a separation or a breaking down. Where two substances are added together to make a third the process is called synthesis, signifying to put together. When there is an exchange of parts between two bodies, as when two substances brought together

WEIGHING

Balances on ring stand. Gram weights and forceps below. Small weights on paper are fractional parts of a gram. Large weights in block run from 1 gram to 100 grams.

give two different substances after the action, the process is known as metathesis, a word signifying "to transpose or exchange." In addition the balance offers a constant means of finding the difference between unlike substances, a comparison we are always making in daily life by the use of our body balance, the sense of touch. We pick up two pieces of cloth of apparently the same thickness and say of one, "This is the heavier." We lift two pieces of wood of equal size and say of one, "This is hard wood. It is heavy." Of the other, "This is soft wood, it is light." So in the study of chemistry we are always judging substances by their comparative weight, saying, "This solid or this liquid is heavier or lighter than water. This gas is heavier or lighter than air"; and by this means we obtain a very real basis of comparison.

Standards. — Whenever we begin to compare we feel at once the need of a standard. The whole world is bound together to-day by a thousand bonds of railroad and steamer. The merchant in New York or London sends his goods to Singapore or Vladivostock, with no more conception that he is doing anything unusual than when he sends them to Boston or to Liverpool. The one thing beyond all else that makes world-wide interchange of commerce possible is the fact that certain standards of weight and measure are fixed, that "a pound is a pound the world around," that a kilogram of sugar is the same weight in Brazil as it is in Paris.

Standardizing weights and measures. — The standards of the world come from London and from Paris. From the first city comes the pound and the yard, the chief units of the English system, from the second the gram and the meter, the chief units of the French system. In those two cities as in the other capitals, buildings are set apart where every possible precaution is taken to keep the standard accurate. Such accurate standards have been transferred from London and Paris to the Bureau of Weights and Measures at Washington, where one finds carefully made standard weights in pounds and kilograms, and bars a yard and a meter long, made with the utmost care and kept where heat or cold, high or low atmospheric pressure, cannot affect them. Every pound weight of sugar that is weighed in this country, every yard measure of cloth that is measured off, was directly or indirectly standardized by those standards at Washington. It is evidently worth while to consider briefly the various measures which we shall use, and to try to find out what relation they bear to one another.

The English and metric systems of measurements. — The English system, the one by which we sell and buy, is but slightly used in scientific work, owing to its unwieldy fractional form, once we go below comparatively large units. It serves very well so long as one does not need to measure below a quarter inch or below a half an ounce, but if our quantities are divided to a thousandth of that amount, it fails woefully. Fractional computations are far harder than decimals, and if you care to turn to the table of weights on the inside cover you will see that grains and pennyweights are too troublesome for ordinary use. But the metric system which we must chiefly employ is simplicity itself because it is based on what is known as the decimal system. We have one decimal system, our coinage, and from that we can readily learn the metric system. The standard of American coinage is the dollar, divide that unit by ten and you get the dime, divide the dime by ten and you get the cent, divide the cent by ten and you get the mill. The dime is one tenth of a dollar, the cent one hundredth, the mill one thousandth. The word dime comes from the Latin decem and means a tenth. The word cent comes from the Latin centum and means a hundredth, while the word mill comes from the Latin mille and means a thousandth. Now take the meter which is the unit of length of the metric system. It is a bar a little longer than a yard, 39.37 English inches to be exact, and was meant to be, and is quite nearly, a ten millionth of the distance from the equator to the north pole. The meter stands as the unit of length in the metric system, as the yard stands as the unit of length in the English system, but where the yard is divided by three to make a foot, and by thirty-six to make an inch, the meter is divided wholly by tens, giving decimals like those into which the dollar is divided. The meter divided by ten gives the decimeter, just as the dollar divided by ten gives the dime. The decimeter divided by ten gives the centimeter, just as the dime divided by ten gives the cent. The centimeter divided by ten gives the millimeter, just as the cent divided by ten gives the mill. In using the measures of length in ordinary scientific work, we take the centimeter as the unit and write a length of two centimeters, four millimeters thus: 2.4 cm. In longer measurements we use the meter, while for very long ones we use the kilometer, which contains a thousand meters and is a little more than six tenths of a mile. That is the measure that you read about in the accounts of automobile and balloon races, and when you hear of a thousand-kilometer race you can know that the contest ran over a little more than six hundred and twenty-one miles.

Measures of volume and capacity. — When the men who made the metric system wanted a square measure which should take the place of our square foot or square inch, they found it in the square meter or square centimeter, which are simple squares, every one of whose sides is a meter or a centimeter long. When they wanted a measure of capacity like our cubic foot or cubic inch, they made a cubic meter or cubic centimeter, cubes each of whose sides is one meter or one centimeter in length. When they wanted a liquid measure, instead of taking our gallon or quart, they took the measures of capacity and used the cubic centimeter and cubic decimeter for the standard. So a cube of liquid, one cm. × one cm. × one cm. is the common laboratory measure for liquids. This is a very small quantity, only about the thirtieth part of an ounce. A thousand times the cubic centimeter is the general commercial measure. This measure (1000 cc.) is called the liter. We are all the time getting foreign commodities measured in liters. Olive oils, chemicals, and alcohol, wines and other liquids come in bottles marked liter, one-half liter, and so on. Most of our measuring of liquids in chemistry, though, is done with cubic centimeters and we write our records in cubic centimeters and tenths of cubic centimeters thus: 2.4 cc.

The gram and the cubic centimeter. — The metric measure of weight, the gram, was taken directly from the cubic centimeter. The men who made the standards had the brilliant conception of taking a cubic centimeter of water at its greatest density, weighing it and making that weight the standard. That brought about one thing of great practical value. One cubic centimeter of water weighs one gram, so we may assume a hundred cubic centimeters would weigh a hundred grams, or, putting it the reverse way, a hundred grams measure a hundred cubic centimeters. As a result, if we want to get any number of grams of water, we can measure out that number of cubic centimeters. If we want any volume of water we can weigh out that number of grams. This fact, that the weight and volume of water are the same numerically, serves us in a thousand places, but it must be remembered that it is *only* with water that this holds true. You cannot, for example, measure out ten cc. of alcohol and think you have ten grams, for alcohol is lighter than water. In no liquid but water does the number of units of weight coincide with the number of units of volume.

The division of the gram. — The gram in its divisions follows exactly the division of the meter. The tenth of a gram is a decigram as the tenth of a meter is a decimeter. The tenth of a decigram is a centigram, and the tenth of the centigram is the milligram, while a thousand grams is a kilogram. We write our results in grams and decimal fractions of a gram. We never write, for example, 5 grams, 3 decigrams, 2 centigrams, 1 milligram, but always 5.321 gm.

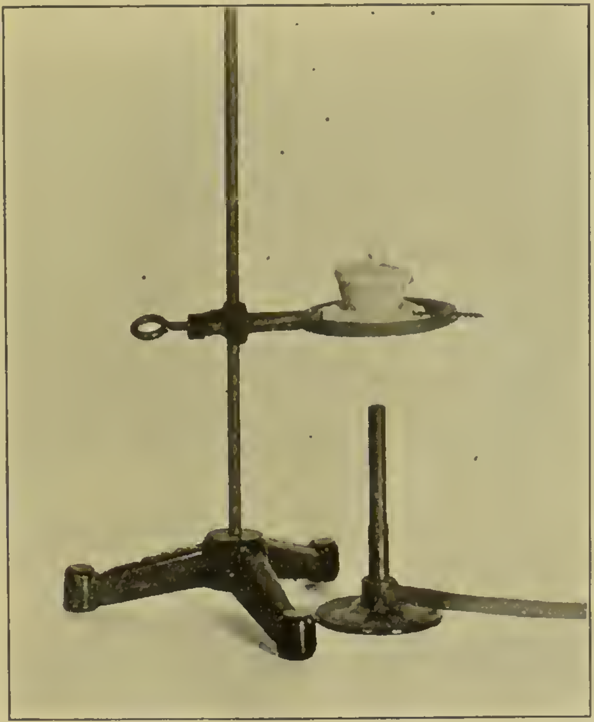
Centigrade and Fahrenheit thermometers. — One other measurement must be recorded before going further, namely, measurement of temperature. We are continually forced to speak of changes which take place at this or that temperature, of bodies which are liquid at one temperature and solid at another. To measure temperature we have two thermometer scales, the Fahrenheit which is found on the thermometer that hangs outside the window in most homes and offices, and the Centigrade which is used chiefly in scientific work. The freezing-point of the Fahrenheit, as we all know, is at 32 degrees. The boiling-point is at 212 degrees. The zero-point of the Fahrenheit is 32 degrees below the freezing-point of water. In the Centigrade system, on the other hand, the freezing-point of water is taken as the zero, while the boiling-point of water is 100 degrees. Contrasting this with the Fahrenheit, it appears that the Fahrenheit freezing-point, 32 degrees, is the Centigrade 0 degrees, while the Fahrenheit boiling-point, 212 degrees, is the Centigrade 100 degrees.

Metrology. — There is a whole science of measurement, the science of metrology of which we have given but the briefest survey in the past few hundred words. But it is essential that we should have this survey of the science in order that when we enter the domain of chemistry we should understand the weights and measures of that domain just as we should know the coinage of the German Empire, know how much our dinner costs or our lodging for the night, provided we are traveling in Germany. The scientific world uses the metric system and all the metric system runs by tens. Its standards for length are the meter and the centimeter. Its standards for capacity are the cubic centimeter and the liter, and for weight the kilogram and the gram. Its measures of temperature are taken with the Centigrade thermometer.

The ratio of weight to volume. Density. So much for the measures of this chemical domain. Now for some of their uses. When we weigh things by the sense of touch and compare one substance with another, the mere weighing alone is not sufficient to tell very much about the matter under discussion. A big bag of feathers may weigh less than a small piece of lead, a large piece of pine may weigh less than a small one of ebony. Therefore there is another factor which we are always forced to consider, the factor of volume. On volume and on weight depend most comparisons. If we have the same volume, the same size piece so to speak, we can tell at once whether one substance is heavier or lighter than another. Then we have a sure touchstone. But unless we use the same volumes, or at least known volumes, we can tell very little about it. Density is the name given to the comparison of the weight and volume of a body, a ratio which may be defined as the amount of matter per unit of volume. We measure amount of matter in pounds or grams, the units of weight; volume by cubic feet or cubic centimeters, the units of volume. If a cube of sugar were two centimeters on every edge its volume would be 2 cm. × 2 cm. × 2 cm. = 8 cubic centimeters. If it weighed sixteen grams its density would be 2, for each of the eight cubic centimeters would weigh two grams if all together weigh sixteen grams. In the metric system density is the number of grams per cubic centimeter, while in the English system it is the number of pounds per cubic foot. Water has a density of *one* in the metric system, for it weighs one gram to one cubic centimeter of volume. In the English system its density is 62.4, for it weighs 62.4 pounds to every cubic foot. Solid volumes of every substance whose density is less than that of water float in water. (We have to say solid volumes here since a hollow iron ship floats because of what is known as its "displacement," but that is another story.) Once we know the density of a substance we know a very valuable property of that substance. Good bricks are separated from poor bricks by their density; good iron from bad iron; good stone from poor. On the density of wood often depends the profit of lumbering operations. If the wood is soft and light it can be floated down the streams at slight expense. If it is hard and heavy it may have to be carried through forests and shipped by train. The freedom from cost of the river transport as against the expense of the road may make a lumbering venture pay or lose. On the varying densities of gases depends the rise of every balloon and airship which floats upward into the firmament, because the gas which it contains is so much lighter than the air around that the lighter body rises in the atmosphere much as pine wood rises in water.

The density of gases. — That question of the different density of gases has much to do with our study of the gaseous elements, since by using what we know of their densities we are able to control and handle them instead of allowing them to escape. You know how hard it is to catch a bird or a dragon-fly. How easily the fluttering thing escapes and keeps before one on an even wing. How much more easily a gas coming from some container can fly to every part of the room. All gases which do not dissolve in water are collected by passing them through water, up into a bottle filled with water. As the light gas rises the heavy liquid falls and the gas takes its place. The gas and water will not mix, and the water is forced out while the gas fills the bottle. If the gas collected is heavier than the air we leave the bottle standing on its base when we test the gas. The heavy gas will stay for a time within as water would stay in a tumbler. If the gas collected is lighter for the time being than air we hold the mouth of the bottle down. The light gas stays inclosed, held up by the layer of heavy air below much as a layer of light kerosene floats on the top of a pail of water. The gas cannot pass through the glass walls of the bottle. It can pass in time through the layer of air, but only slowly, and before it can so pass we shall find amply sufficient time to test it. If we collect a light gas by running it up into a bottle of air, forcing the heavier air down and out, we may say we are collecting it by upward dry displacement. If we collect a gas insoluble in water by bubbling it into a bottle full of water and forcing the water out we call the process collecting over water.

How chemical changes are brought about. — To bring about chemical changes we use various means. Sometimes heat will effect the change, as when coal burns and turns to ashes which are left behind, and to gases which go off in the air. Sometimes light is the agent used, as when the sun's rays act upon the chemicals of a photographic plate and leave a picture there. Sometimes electricity works the change, as when the current dissociates the silver from a silver bath, and coats a fork or knife with silver plating. Sometimes the action occurs by simply bringing two chemicals together, and we may say in general that changes come from separation or from union of substances because of their so-called affinity or chemical attraction for each other aided by light, heat, or electricity, by one or two or all of them.



HEATING A CRUCIBLE

One of our ordinary ways of bringing about a chemical change in the laboratory is the heating of a substance in a porcelain crucible. The crucible rests on a pipe clay triangle placed on the ring of a ring stand. It is exposed to the direct heat of the flame. Flasks, beakers, and evaporating dishes, on the other hand, are never exposed to the direct flame. They are heated on asbestos or on wire gauze.

CHAPTER III

OXYGEN

Symbol for oxygen O. For discussion of symbols and reaction writing see Chapters XIII and XXIII.

The oxygen of the atmosphere. — Wherever a fire burns, whether it be the driftwood fire on the beach or the roaring blazing mass of red coals under the boilers of an ocean liner, the gas oxygen plays the chief part in the burning. Whatever fuel we use, whether it be the wood of a match, the gas of the

A Great City Fire

The wood, like the other combustible matters, turns to oxides, combining with the oxygen of the air. Smoke, largely uncombined carbon, is rushing off in great quantities.

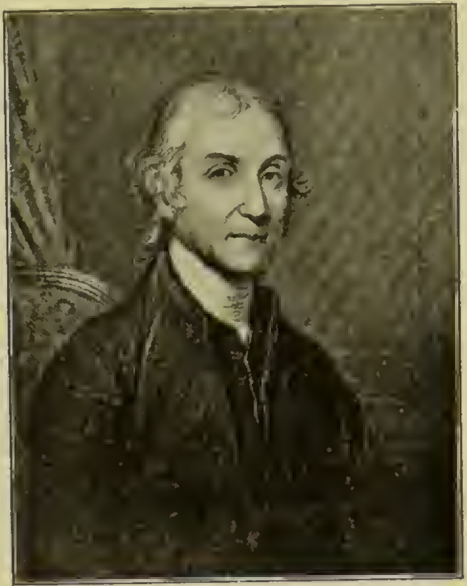
gas-stove, or the kerosene of the study lamp, the heat and flame which is produced comes from the power which the fuel possesses to combine with oxygen; the more oxygen you have the hotter and brighter the combustion or burning. Every time that we open the drafts of a stove or furnace we do it to increase the fire by pulling more oxygen across the hot fuel. Every time we throw water or a blanket over burning material to quench the flames we smother the fire by keeping oxygen from it. Every time we breathe we are drawing life-giving oxygen into our lungs to furnish us with energy and warmth. Since we are wholly dependent on fire of one sort or another for life itself, it is evident that there is some practically unlimited source of supply of this gas oxygen which keeps the fires of all the world alight. This unlimited supply is found in the air about us, the surrounding atmosphere in which we live and move and have our being.

The nitrogen of the atmosphere. — The atmosphere is by no means wholly oxygen. If it were, as we shall see, the whole world would soon blaze up, a mass of flames. About one fifth of the air is oxygen, and the other four fifths is chiefly composed of another gas, nitrogen, which serves to keep the oxygen in subjection. The properties of this gas, nitrogen, will be considered later, but its relation to oxygen is the one which interests us now and is that just stated, the weakening or diluting of the force of this gas which is necessary to fire.

What is burning? — We say that oxygen is the cause of ordinary burning; that brings us to a question which troubled many wise men for two or three centuries. What is burning? What is the nature of flame? We put wood or coal in a kitchen stove and it burns, giving off heat by which we cook food or warm ourselves. After the burning there seem to be left only a few ashes whose weight is but slightly comparable to the weight of wood or coal with which we started. This happens with the ordinary fuel of the stove. On the other hand, if we burn metals such as iron or mercury, the weight of the substance left after burning is greater than the weight of the substances which we had in the beginning. Here is a contradiction. Certain substances when burned apparently gain in weight. Certain other substances when burned apparently lose in weight. What really happens? The answer once known seems easy, but its solution tasked the best efforts of the whole scientific world for a couple of hundred years. When wood or coal is burned, matter in two states results. A solid part is left, the visible ash. A far heavier gaseous part goes off through the chimney into the air. That is the secret of the fire. The hot gases which rise are not the visible smoke; they are the invisible gases which rise from both a clear fire and a smoky one. All substances when burned with oxygen gain in weight. The products from the burning of wood or coal, iron or mercury, or whatever the substance burned may be, are heavier than the things burned. When metals are burned we can see that they are heavier by weighing the solids left after burning. When wood or coal is burned we find the products of burning are heavier by weighing both the ashes left and the gas which comes off.

The compounds of oxygen. — This statement leads us to the next step in our discussion of the fire. If all things burned become heavier, something must be added to them. Since oxygen is all about us in the air, and is essential to burning, it is not hard to see that the substance which is added is this same gas. Ordinary burning is really the combining of oxygen with some other substance to form a compound of oxygen called an oxide. There is the explanation of all the puzzles. If burning is combining with oxygen (oxidation is the name of the process), then the more oxygen you add to any burning substance the quicker the substance will burn, the hotter the flame. That is why you open the drafts of the stove. When substances are oxidized, oxygen has been made a part of their composition, and they must weigh more than they did at first. The ashes which are left in the ash pit are largely solid compounds of oxygen, the gases which go off are gaseous oxides.

The discovery of oxygen. — It was not until 1774, just before the American Revolution, that any one knew why a fire burned or that there was such a gas as oxygen. In that year an Englishman, Joseph Priestley, took mercury (the liquid metal inside the thermometer tubes which you know as quicksilver), and burned it in the air. It turned to what was known as red ash or red powder of mercury. This substance was a red soft powder of comparatively light density, not unlike flour in texture and wholly different from the dense shining silvery liquid from which it had come. The experiment so far was no different from what had been done before, without result, but Priestley, with that brilliant imagination which has so often characterized the great leaders of science, saw a new possibility. If mercury had changed to a red ash by burning, could not the substance which had so changed it be obtained from the red ash in its original form by reheating? Priestley tried the experiment by taking a little red powder in a tube and heating it in such a way that any gas which escaped could be collected over water. Soon clear, colorless, odorless bubbles of gas rose and filled the receiver. The liquid mercury with which the experimenter started clustered in drops on the side of the tube. The red ash disappeared and mercury was left in the tube, while a gas escaped and filled the receiver — an unknown gas whose bubbles looked in all respects like bubbles of air. When a glowing splinter was plunged into the bottle full of this strange gas, the action was different from that which would have occurred had the test



JOSEPH PRIESTLEY. (1733-1804)
Discoverer of Oxygen.

been made in the atmosphere. In air the splinter would have glowed dully for a minute and then gone out. In this gas it relighted and burned brilliantly. All substances burned far more brilliantly in this substance than in air. The slightest glow kindled into brilliant flame, while the gas united with intense eagerness with almost every substance that was tried. The gas was the element oxygen. The secret of combustion was explained. Ordinary burning was proved to be a process of combination with oxygen. Burning was far swifter in oxygen than in ordinary air, because air contained only one part in five of oxygen, the other four parts being nitrogen, the gas which stops combustion instead of aiding it.

Two ways of making oxygen. — To this day oxygen is prepared in the laboratory from the same material which Priestley used, the compound of mercury and oxygen, red oxide of mercury. To gain a greater amount of oxygen, however, we use a quicker method, and prepare it from a mixture of potassium chlorate and manganese dioxide, two chemical compounds, whose chief interest to us lies in their power of producing this substance.

Writing an equation. — Such a chemical change as Priestley brought about, whether produced by addition, by division or by interchange of parts is commonly called a *reaction,* and can be expressed in the form of a general equation. Mercuric oxide when heated breaks down into mercury and oxygen, and the equation or reaction which expresses this may be written thus:

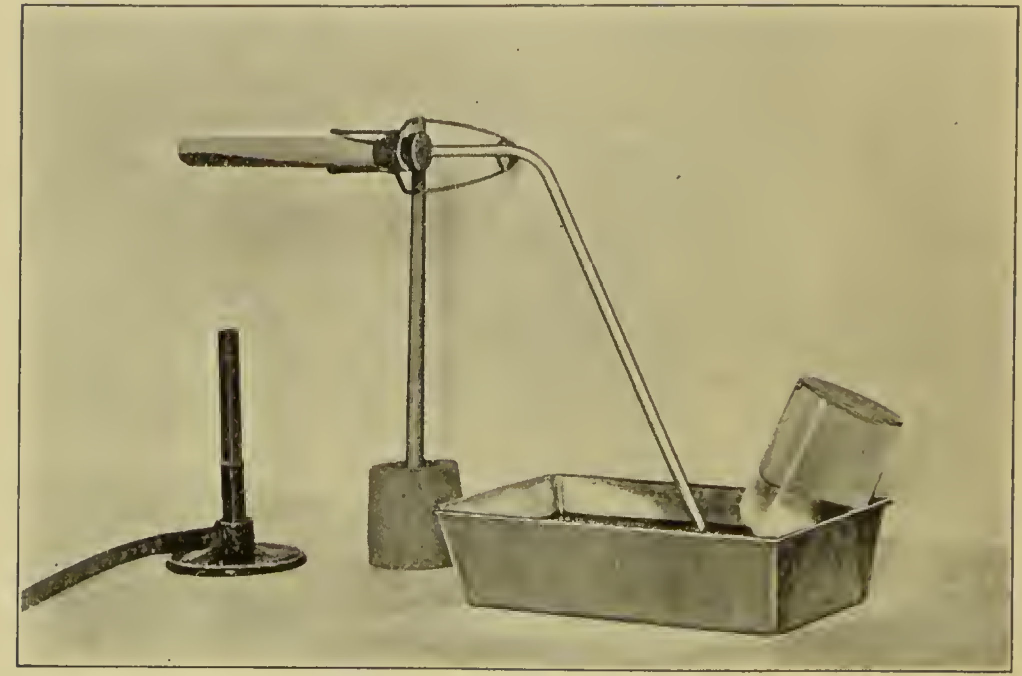
Mercuric oxide (heated) = Mercury + oxygen. HgO (heated) $= Hg + O$. Such an equation would express the action briefly and clearly.

Note. — These equations, written in symbols, which appear here for the first time, may be disregarded until the student has mastered Chapter XIII on equations. After equation writing has become somewhat familiar, it will be well to turn back and review all the equations inserted in the text. After three months' study on descriptive chemistry, the problems of equation writing should become comparatively simple and should yield excellent results in the way of making clear the various changes which take place.

Catalytic action. — The white crystals of potassium chlorate contain much more oxygen than does the red oxide of mercury. These crystals give off this gas fairly readily when heated by themselves, but do this far more readily when the black powder, manganese dioxide, is mixed with them. In this process we find a strange and rather inexplicable phenomena. The manganese dioxide gives off no oxygen, suffers no change that we can recognize, yet its presence enables the potassium chlorate to send off oxygen far more readily. It is as if a man were turning off a certain quantity of work each hour on a dark cloudy day. Do what he will he cannot pass beyond a certain limit. The sun comes out and floods the room with cheerful radiance. The man's hands work faster, — swifter and swifter grow his motions. In the sunlit hours he passes far beyond the utmost limit of work accomplished in the dark hours. The sun has done none of the work. It has helped the worker. So the manganese dioxide has helped the potassium chlorate to give off its oxygen without apparently being permanently affected itself. An action such as this, where a substance which assists a chemical change comes out of it seemingly unaffected and without the slightest apparent change in form, is called a catalytic action or catalysis. Some of the most important advances of modern chemistry have been accomplished by the aid of catalysis. As the manganese dioxide acts as a catalytic agent we do not include it in the reaction which we write for the preparation of oxygen, and simply say Potassium chlorate = potassium chloride + oxygen.

$$KClO_3 = KCl + 3O$$

Collecting a gas over water. Properties of oxygen and a test for the gas. — The oxygen which flows rapidly from a heated mixture of potassium chlorate and manganese dioxide is collected over water. This element is very slightly soluble in water, so that the bubbles pass up into a water-filled bottle, driving down the water which the bottle contains, and filling the space with gas. To understand the way in which bubbles of gas will force water out of a bottle, think of what happens when you try to pour a liquid out of a full bottle with a narrow neck. If you turn the bottle straight upside down, oftentimes the liquid will not come out at all, and if it does come it flows with difficulty. If you turn the bottle on its side so that air may enter, big bubbles of air enter through the upper part of the neck, and by just so much as the air enters, by just so much does the liquid flow. The air entering forces the liquid out and takes its place. We see the bubbles of air forcing their way in, and yet air is invisible and colorless. We see the bubbles by their contrast with the liquid within just as we see the colorless bubbles of the gas in soda water forcing their way to the top of the glass. Like air oxygen is colorless, odorless, and tasteless. Like air it forces water out of a bottle. It is slightly heavier than air, a fact which makes us keep it



PREPARATION OF OXYGEN

A mixture of potassium chlorate and manganese dioxide is being heated in a hard glass tube. The gas passes out through a bent delivery tube into the receiving bottle placed in a pneumatic trough (here a cake tin) and drives out the water.

in bottles with the mouth up, and it is tested by thrusting a glowing splinter into the bottle holding the gas. Such a splinter in a bottle full of oxygen instantly relights and burns brilliantly. The low blue flame with which sulphur, such as the sulphur of a sulphur match, burns in the air turns to a high brilliant blue in pure oxygen, and iron tipped with sulphur burns in the gas with brilliant sparks.

Oxides in nature. — The compounds of oxygen, the oxides, are of almost as much importance as oxygen itself. Since oxides are left after all ordinary burning, it may be imagined that they are found plentifully in nature. The belief that this earth was once a fiery molten mass, finds basis in the fact that most of the earth's crust, rocks and sand, dry earth and marsh, is made up of oxides or other compounds of oxygen. Many of the metals are found in the mines in the form of oxides so that the oxygen has to be removed from the ores in order to get them in metallic form.

The preparation of oxides. — To prepare oxides it is generally necessary to heat a substance in oxygen or in air. Most of the metals oxidize in this fashion, as copper burning changes to copper oxide, or as the iron rust which we see daily shows that iron has burned. The carbon part of wood, coal, oil, or gas burns producing carbon dioxide (the prefix di, meaning two, shows that carbon unites with two atoms of oxygen) CO_2. Sulphur burns in oxygen, forming sulphur dioxide, SO_2, and hydrogen burning in air forms hydrogen oxide, H_2O. In fact, there are only seven common elements which do not unite directly with oxygen, and only two, bromine and fluorine, which do not unite with it at all. The oxides appear in so many different forms that no general tests for their class can be given. The more important have specific tests which will be taken up later.

Different cases of oxidation. — No chemical process has greater value than has oxidation, and probably the most valuable single case is found in the formation of one gas, carbon dioxide. When carbon is burned in oxygen, heat is given off.

The coal burning in the furnace and the food burning in the body are both examples of heat produced by oxidation. These are comparatively rapid processes as are many cases of ordinary oxidation, but sometimes this action is so slow that it is recognized only by its effects. Iron left in the air for months turns slowly to iron rust. During the whole period of its exposure, night and day, winter and summer, it has been burning, and the oxygen of the air has been slowly combining with the metal and forming iron oxide, and another oxygen compound, iron hydroxide. This oxidation gives off as much heat and is as complete as when the same amount of iron, blazing into a thousand sparks in an atmosphere of oxygen, changes into iron oxide in five minutes. The one is slow combustion. The other is swift combustion. Paint drying on the house wall dries, in part, because the oil which rendered it liquid and enabled it to spread has oxidized at the temperature of the air. Paints are made by mixing colored solid substances, such as white lead, with oil. The combination makes a liquid easily spread on wall of house or barn. Once spread where air and sunshine can reach it a part of the oil of the mixture oxidizes, the gaseous oxides go off and the solids are left behind coating the wall. Vinegar comes from cider because tiny plants of "mother of vinegar" have enabled the oxygen of the air to be added to the alcohol of the cider, and so made it possible for the alcohol to change to vinegar.

Oxidation in the human body. Respiration. More nearly concerning us is the use of oxidation in the processes of our daily life. The human body is a cunningly devised engine using food for fuel. The burning of that food in the body produces the heat necessary for the body warmth, and the energy necessary for all thought and action. For that burning, oxygen is necessary, and with every breath the lungs take in oxygen from the air. The burned wastes of the body, the oxides produced by the combustion, escape through the lungs and pass out with every breath. The taking in of air, bearing oxygen for body burning, and the giving out through the breath of the oxides which remain after burning, is called respiration. On respiration depends life itself.

Uses of various oxides. — The use of oxides in fireproofing materials and in extinguishers has grown largely in late years. Substances once burned cannot be burned again, and a wall made of oxides can stand a flame without injury unless the heat expands it. Buildings are made fireproof by using in their construction substances which will not burn, such as stone, brick, iron, cement, concrete, and asbestos. Most of these substances either do not burn save at very high temperatures, or have taken up all the oxygen which they are capable of holding, and, since they can take no more, burning becomes impossible, since burning means the addition of, or combination with, oxygen. Clothes can be made fireproof by saturating them with incombustible materials. Wood can be soaked in liquids of this sort which dry in it, protecting not only the outside surface, but filling the whole substance of the wood. Fireproofing of carbon in any form, however, only protects it from fierce flame such as conflagration. It cannot prevent some smoldering. The use of carbon dioxide in fire extinguishers will be considered later.

How oxygen combines. — To bring about the union of oxygen with any other substance commonly requires greater heat than is found in ordinary temperatures. Phosphorus, however, burns at low temperatures, and phosphorus oxide forms very easily. Hence came the phosphorus match, easily ignited by friction. That friction gives heat is seen every time a saw cuts a board. The phosphorus of the match takes fire first, and then kindles the sulphur of the match, which, in turn, kindles the wood of the match. In a kitchen fire a match kindles paper, the paper kindles wood, and the wood the coal. So we have here a whole series of oxides formed each with a different kindling point, or temperature at which oxidation takes place.

The uses of oxygen. — Oxygen plays more parts in our daily life than any other element, is found more widely, and makes more compounds; through its union with carbon giving off heat it keeps trolley car and mill wheel turning, and gives the light of glaring arc, and study lamp. Most important of all it is essential to the processes on which human life depends.

Ozone, the allotrope of oxygen. — Beside the ordinary kinds of oxygen we have another, an unusual type, ozone, concerning which an illustration used by one of the ablest of British chemists has never been surpassed. "Ozone," he said, "is the alias of oxygen." Suppose we carry out his simile. When Jones, alias Brown, is tried in court for burglary, we know very well that Jones has been plying his criminal trade in one place under his own name, and in another place under the name of Brown. The two names stand for the same man. Sometimes under the name of Brown he wears a beard, while under that of Jones he is clean-shaven. The outward appearances of Jones and Brown may differ, but the man himself is unchanged. So oxygen appears under another form and has an alias, ozone. In chemistry, when we find an element which, though it remains an element, yet appears under more than one distinct and different form, we say concerning it that it has the property of allotropy, while we call the different shapes in which it appears allotropic modifications or allotropes. Besides oxygen, which has the allotrope of ozone, yellow phosphorus has the allotrope of red phosphorus, while carbon appears as diamond, graphite, and amorphous carbon, and sulphur appears in four allotropic forms.

The production of ozone. — Ozone, the allotrope of oxygen, while composed wholly of the element oxygen, is strikingly different from the usual form. It is commonly made by passing oxygen through a tube under the influence of what is known as the silent discharge of electricity, that is, sending the gas between a host of tiny sparks which spring from one electrically charged wire inside the tube to another wire wound round the tube. When oxygen is exposed to such an electrical discharge, ozone, its allotrope, is formed. The symbol for a molecule of oxygen = O_2. This symbol represents two atoms of oxygen. The symbol for a molecule of ozone = O_3. This symbol represents three atoms of oxygen. What really happens in this formation of oxygen's allotrope is that the oxygen molecule, which normally contains two atoms, under the influence of this discharge recombines and forms a molecule (ozone) containing three atoms. To this change in molecular structure the differences between the two forms are due.

A test for ozone. — Oxygen does not readily attack organic substances like rubber. Ozone attacks such substances, destroying rubber tubing rapidly. Oxygen does not affect such metals as silver and mercury. Ozone combines with them swiftly. The medicine iodine, which is used to paint some kinds of body injuries, combines with the metal potassium to form a substance known as potassium iodide. In this shape it has no effect on common starch. Free, it colors starch blue. Expose potassium iodide and starch to the action of oxygen and nothing happens. Expose the same substance to the action of ozone and the starch turns blue immediately. The ozone has torn the potassium from the iodine, whereupon the latter substance, suddenly liberated, turns the starch blue. That is the test by which we can recognize the gas in its pure state. Ozone is seldom permanent. It turns back to oxygen when heated, releasing an atom of oxygen as it breaks down. This loss may be represented as follows:

$$O_3 = O + O_2.$$

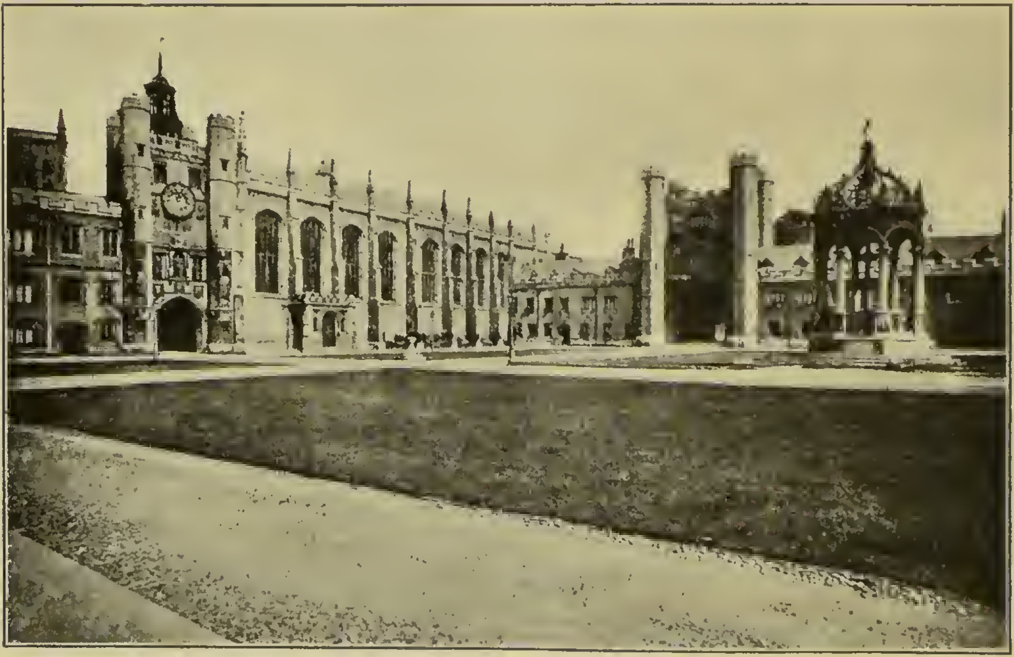
Despite this fact it is found to some extent in fresh country air. It is formed in the atmosphere by the electrical discharge of the mighty generators of the heavens, made by the forked lightning as it flashes across the sky.

CHAPTER IV

HYDROGEN

Symbol for hydrogen, H

Cavendish and the discovery that burning hydrogen forms water. — Of all the leaders of English thought that the gray old halls of Cambridge University have sent into the world, few were at once as eccentric and as clear-visioned as Henry Cavendish. His birth in 1731 found the chemistry of the gases a dark and confused chaos, concerning which philosophers disputed vainly. His death in 1810 found the chemistry of the gases in its first essentials much the same field that it is to-day. That change was largely due to Cavendish's clear



Trinity College, Cambridge, England

At Cambridge University Henry Cavendish and many other British Scientists were educated.

reasoning, accurate experimentation, and complete observation. None of his brilliant researches were more remarkable than those on hydrogen. He was the first scientist to see that the hot, pale, blue flame of hydrogen, as it rose from a jet, was burning in oxygen, and that, as the product of action between this burning inflammable gas and the life-giving oxygen of the air, water was formed. By that historic study Cavendish conclusively settled questions which had troubled scientists for two centuries and demonstrated that the answer to the question of "How is water formed?" is, "It is formed by the burning of hydrogen in oxygen." To the question "What is water?" Cavendish gave the reply, "It is hydrogen oxide."

Occurrence of hydrogen. — Found as an element in but a few rare substances, in meteorites for example, hydrogen oxide covers more than half the world; for when hydrogen burns in oxygen, hydrogen oxide is formed and hydrogen oxide is water, the water of the sea and the river, of well and brook. The whole great class of the acids, from sulphuric acid (the oil of vitriol of commerce) to the sour citric acid of the lemon, owe a large part of their qualities to hydrogen. That element is the essential part which makes the substance acid.

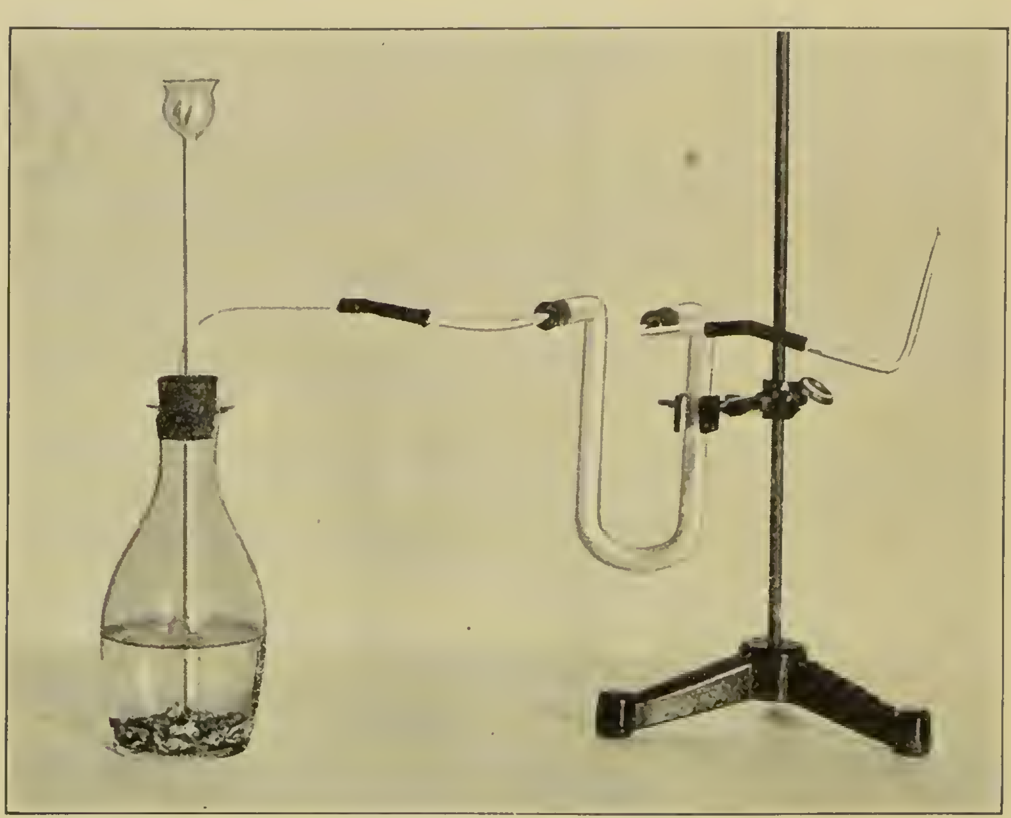
The weight of hydrogen. — Lightest of all the gases, the weight of an atom of hydrogen is the chemical standard of weight for the other elementary atoms. The reason why hydrogen is taken as the standard of weight is chiefly because of the fact just stated, that it is the lightest of all gases. The weight of any gas may seem to be too small to be a standard of value, but when we realize that the weight of air in a small room 10 ft. × 20 ft. × 10 ft. amounts to 160 pounds we can see that on varying weights of gases many phenomena may depend. Watch a child's balloon rising into the air. Its ascent depends solely on the fact that the gas inclosed in the round shell is so much lighter than the air around it that it pulls up the whole mass. A piece of wood rises in water for the same reason that a balloon rises in air, because its density is less than that of the water around it. A hot air balloon sent up on a Fourth of July night rises because the air within the balloon bag, heated by a flame, expands, grows less in weight, and so pulls the light envelope up into the summer sky. The weight of gases indeed may vary almost as widely as the weight of solids. The air is light, but hydrogen is less than one fourteenth as heavy as air.

Using a gas for measurement. — Right here a question rises. How can we use hydrogen gas for measuring? We can put a pound weight on a scale, pour a quart of milk from a tin measure, or lay off ten feet on a board with a rule, but here is a wholly different proposition. Here is a gas which will not stay in any one place unless carefully confined. Think a moment and the explanation appears. A pound weight is simply taken from the standard pound, as we already know. If we once weigh a liter of hydrogen correctly (an experiment we can accomplish by weighing a bottle full of hydrogen and weighing it practically wholly empty, that is, when it contains neither hydrogen, air, nor any other gas), we have a standard which will serve us permanently. A liter of hydrogen has been weighed carefully, at a given temperature and pressure, a great many times, and the weight that we accept as the standard for the liter is .090 grams (formerly taken as .0896). That is, if we take a thousand cubic centimeters of this gas it weighs a little less than a gram, than a thirtieth part of an ounce. Knowing the weight of that volume of hydrogen we can say that air, for example, is 14.4 times as heavy as hydrogen, that oxygen is 16 times as heavy as hydrogen, or that chlorine is 35.5 times as heavy as hydrogen. We have by this means a standard, formed just as we formed the standard of the yard and of the meter. If we say that a gas is heavier than air we know something about it, but nothing very definite. If we say that it is twenty times as heavy as hydrogen, we know how heavy it is, know not only that it weighs 1.8 grs. per liter compared with hydrogen's weight of .09, but also that it is about a fourth heavier than either air or oxygen, which weighs, as we have said, 14.4 times and 16 times respectively as much as hydrogen.

A general method for making hydrogen. — So much for lightness. Now as to the production of the gas. Evidently, since hydrogen is in all acids, these substances should form a field from which we can obtain it. All acids give up their hydrogen when properly treated, and the easiest way to obtain the gas is by acting on some strong acid with a metal. We can use a whole series of acids and a whole series of metals, but zinc, the zinc which lines the water-tank, acts most easily and cheaply as the metal, while the acid that we use may be either sulphuric or hydrochloric as both give up their hydrogen readily when brought in contact with zinc.

A special method for making hydrogen. — To make hydrogen, all we have to do is to place some zinc in a bottle, furnish a tube through which we can pour acid down, and a tube through which the hydrogen can escape up. As soon as we pour the dilute acid on the zinc, the bubbles of gas begin to rise, for the zinc forces the hydrogen out of the acid and takes its place. The hydrogen which comes off is invisible and odorless, its weight is less than that of air, and it is insoluble. On account of that property of insolubility we can collect it over water as we did oxygen. On account of its lightness we can collect hydrogen by simply holding a bottle mouth down over the end of the delivery tube through which the gas escapes. The light gas rising forces the heavy air down just as it forces the water from the filled bottle on the trough. When any gas is materially lighter than air, we can collect it in this way, a method called upward dry displacement.

The action in a hydrogen generator. — All the time inside the hydrogen generator, while the tiny bubbles rise through the acid and churn the whole into a froth, a wonderful action is going on. Imagine a whole army corps of soldiers uniformed in red and ranged by twos, each pair of men linked together by a band of steel. Now across the field whereon they stand suppose another corps clad in black in single rank comes charging on them. Just as they reach the red line every black man of the second force leaps at one banded red couple, breaks the chain, forces the red man on the right backward with a sudden blow and takes his empty place. As each black man takes one red man's place the chain suddenly welds itself together and the corps again presents a long line of pairs. But not as it did before. Every pair now is particolored, one man black, the other red. The second corps now leaving the field in single rank is red where the oncoming host was black. The whole



Preparation of Hydrogen. I.

The gas is generated in a thick glass bottle fitted with a thistle tube and a delivery tube. The thistle tube dips below the surface of the liquid. The delivery tube reaches only to the bottom of the stopper. The gas may either be collected by displacement over water as oxygen was collected, or by upward dry displacement as here. In this latter case a calcium chloride drying tube removes water present in the gas This removal makes it evident that water produced by burning hydrogen comes from the burning of the gas itself and not from water carried over by the gas.

condition of affairs has been reversed. Such an action takes place when zinc attacks sulphuric or hydrochloric acid. We may consider that both of these acids are made up of two distinct parts. The hydrogen is one part in each acid, while a little group of atoms which for convenience' sake we shall call the sulphate radical (we shall discuss and define this term later) makes the second part of sulphuric acid, while an atom of chlorine makes the second part of hydrochloric acid. The hydrogen and the sulphate radical in the first case are bound as were the soldiers of the army of the reds. Hydrogen and chlorine in the second case are bound in similar fashion. The zinc comes charging on them like the army of the blacks, drives off the hydrogen with a swift blow and takes its place. An atom of zinc together with the group of atoms (the sulphate radical), or else an atom of zinc and an atom of chlorine, are left united, hydrogen goes free and seeks to escape through the frothing mass of bubbles in the generator. Writing these as equations we may say:

Zinc + hydrogen and chlorine = zinc and chlorine + hydrogen.

$$Zn + 2HCl = ZnCl_2 + H_2$$

Zinc + hydrogen and the sulphate radical = zinc and the sulphate radical + hydrogen.

$$Zn + H_2SO_4 = ZnSO_4 + H_2$$

The gas hydrogen escapes in either case. A new compound made up of zinc and the sulphate radical, or else of zinc and chlorine, is left in the generator.

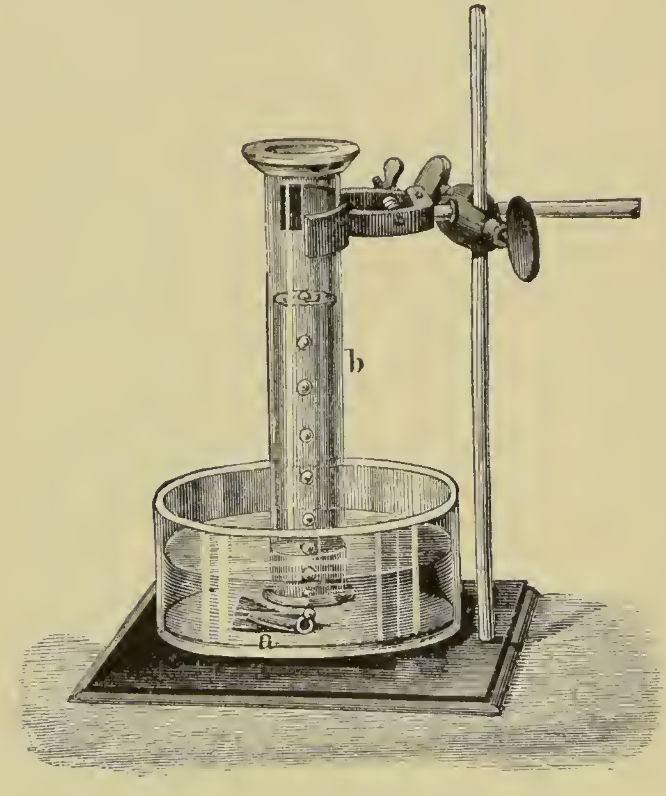
Chemical solution. — The action of sulphuric acid on zinc with the formation of zinc sulphate (as the compound is called which is formed from zinc and the sulphate radical) has another side beside the direct formation of hydrogen gas. It is an example of what is known as chemical solution. When you dissolve sugar in water the sugar remains unchanged in form in the solution. If we separate the water all the sugar will be found at the bottom of the vessel. This is a physical change. Such solution is called physical solution. Here we dissolved the metal zinc in sulphuric acid and formed a new compound, zinc sulphate. When we evaporate the solution we have left zinc sulphate, not the zinc with which we started. When a chemical change takes place under such circumstances of solution the action is called chemical solution.

Testing hydrogen. — Suppose this action between metal and acid has been accomplished, and we have the hydrogen safe in the bottle held to receive it. We must hold the bottle mouth down so that the light gas will not escape. Then we can try the flame test we tried with oxygen. But before bringing a flame near hydrogen a word of warning must be spoken.

Hydrogen explosions. — In its *explosive* power lies the danger of experimenting with hydrogen. *Every care must be taken to free* the hydrogen from air before it is lighted, to see *that no flame* is brought near a hydrogen generator when there is any chance that hydrogen may be escaping. The explosive power which may be produced when a mixture of hydrogen and oxygen is lighted is sufficient to burst a glass generator or receiver and bring on serious results. There is *no danger* if a few simple precautions are taken; if, for example, the hydrogen is run for a few minutes before testing, so that the air which fills the generator may escape and hydrogen may take its place, and if the hydrogen to be tested is *taken away* from the generator before lighting and no flame is brought *near* the generator. If, after taking these precautions, we now take a lighted match and touch it to the gas at the mouth of the bottle where the hydrogen meets the layer of oxygen carrying air, a slight explosion generally occurs and a blue flame darts up into the bottle. There is the test for hydrogen. It burns with a blue flame when it unites with pure oxygen or with the oxygen of the air. Remember if a serious explosion occurs it is because the hydrogen in the bottle was mixed with oxygen, and be sure you do not put a bottle in which hydrogen is still burning over the generator tube. One other point should be noticed here, if we thrust a glowing splinter swiftly into the gas it goes out. Evidently the gas itself will not support combustion, will not allow burning.

Sodium and potassium. Their use in preparing hydrogen. — One other method of preparing hydrogen is of such general importance that it cannot be omitted here. It has to do with two elements, sodium and potassium, the first the metal found in common salt, and the second, the metal found in common potash. Both of these substances are wholly unlike any of the metals we ordinarily meet. A metal to us means a dense substance, which sinks in water. Both of these elements float high above water like light wood, borne up by the light gases which rush from their union with the water. Most metals are hard. These two are soft like putty. Most metals are quite unaffected by liquid water. Drop either of these substances into water and a most energetic action takes place. Bubbles of gas rise swiftly to the surface of the water, the metals race around on the surface and so much heat is generated in the action that the gas generated by the sodium and potassium takes fire and burns if the action is confined to a single spot. Yet despite all these variations from the usual behavior of their class, sodium and potassium are true metals in all the essential things that characterize the group. They conduct heat and electricity, have a metallic luster and form metallic compounds.

The hydrogen of water. — The action of these metals on water illustrates in a striking fashion the fact that hydrogen is one of the constituent parts of water. The action of both metals is very similar, so we may take the action of sodium as an example. If we place sodium in a cage of wire gauze so that it cannot escape, but must act in a single place, and then drop it into water, big bubbles of hydrogen gas rise through the liquid. They may be collected in a water-filled bottle. The hydrogen cannot come from the metal sodium which is an element. It must come from the water, which is a compound containing hydrogen. Though water is hydrogen oxide, the substance left behind is not sodium oxide. Only a part of the hydrogen goes off as a gas, one atom of the hydrogen stays behind and unites with one atom of sodium and one atom of oxygen to form a new compound, sodium hydroxide, which we shall study more carefully when we come to take up acids, bases, and salts. It is enough for our purpose here to know that when either sodium or potassium is dropped on water, hydrogen goes off and sodium or potassium hydroxoxide is left behind.



Preparation of Hydrogen. II.

Production of hydrogen by the action of sodium on water. The sodium held below the water's surface at *a* sets free hydrogen gas, which rises rapidly into the jar *b*. *Use great care here.* Do not use a piece of sodium or potassium larger than a pea. Do not leave either metal exposed to the air. Do not touch them with bare hands.

$$Na + H_2O = NaOH + H.$$

$$K + H_2O = KOH + H$$

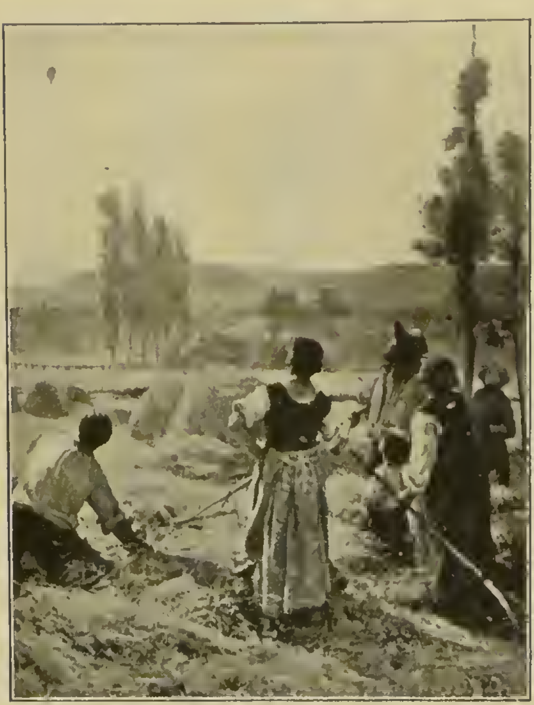
Water. Hydrogen oxide. — When Cavendish saw hydrogen burning in a dry clear bottle he noted small drops of water after the flame had passed. We can repeat that experiment to-day, and, with our knowledge already gained that the result of burning substance in air is the formation of oxides, expect burning hydrogen in air to produce hydrogen oxide. That the result of such burning should prove that hydrogen oxide is water, the ordinary water that we drink, the water of the rain and of the sea, is a matter quite beyond all ordinary expectation. Here we have a light gas burning with a blue flame, the product of whose burning is water, the liquid that we use to put out flame. It is one of the marvelously simple contradictions of chemistry.

Burning hydrogen. — The burning of hydrogen with the actual production of water may be accomplished readily enough by lighting a small tip of glass attached to the generator, but trial after trial should be made until the gas gives but the slightest sound on ignition, as it is dangerous to ignite it while any air is present. Once lighted, the pale blue flame gives a tremendous heat. Refractory metals give way before the roaring flame of the oxyhydrogen blow pipe in which streams of pure hydrogen and pure oxygen meet in flame. The lime light of the stereopticon on the stage comes from lime heated to incandescence with the heat of such a flame. Only by the electric furnace is it surpassed in heating power. The reaction for burning hydrogen is $2H_2 + O_2 = 2H_2O$.

And now one point. Reaction writing may be carried on in either one of two ways. The equation just given is often written in its simplest form, $2H + O = H_2O$. It is quite as commonly written in the way just cited, $2H_2 + O_2 = 2H_2O$.

The second form is preferred by many as it indicates that the two atoms of the hydrogen and oxygen molecules take part in the reaction which goes on. To illustrate the principle of equation writing by the second method, let us consider what is meant by the expression ($2H_2$) in the equation above. The first figure 2 written before the H and on the line indicates the multiplication of two molecules of hydrogen. The second figure 2 written after the H and below the line indicates the number of atoms in the hydrogen molecule. In general a number placed before a symbol or a formula indicates the number of molecules of the substance named which enter into the reaction. A number placed after a symbol or a radical indicates the number of atoms or of groups of atoms present in the molecule represented.

The lightness of hydrogen. — The lightness of hydrogen has made it of use in many of the air-ships which have been made in recent years. No other gas has its lifting power. As a soft pine plank will bear a greater weight in water than a hard chestnut plank of the same size because of its lesser density, so hydrogen, having the smallest density of any gas, possesses the greatest lifting power to raise the ships which ride among the clouds.



THE LIGHTNESS OF HYDROGEN

A balloon rising because its hydrogen-filled sphere is lighter than the air which surrounds it.

Preparing hydrogen from steam. — The properties of water, the chief compound of hydrogen, are so many and various that they must be taken up in a separate chapter, but one way in which hydrogen can be obtained from water should be mentioned here. When steam (gaseous water) is passsed over red-hot metals or carbon, it breaks down into its component parts, the oxygen unites with the metals or carbon, forming oxides, and the hydrogen goes free. For example, if we pass steam over red-hot iron we could write an equation thus:

Hydrogen oxide + iron = iron oxide + hydrogen.

$$3Fe + 4H_2O = Fe_3O_4 + 4H_2.$$

The passage of steam over red-hot iron is sometimes used to obtain hydrogen in the laboratory. When steam is passed over red-hot carbon the gases resulting are used, as we shall see later, to form the "water gas" which lights many of our streets.

Hydrogen dioxide. — Water is not the only oxide of hydrogen. There is a second oxide, hydrogen dioxide or peroxide, which contains two atoms of hydrogen to two atoms of oxygen. This oxide is a liquid and is commonly used in a solution containing somewhat less than five per cent of the dioxide to ninety-five per cent of water. It breaks down easily into water and oxygen as the following equation shows:

Hydrogen dioxide = hydrogen oxide (water) + oxygen.

$$H_2O_2 = H_2O + O.$$

The oxygen which is thus freed serves a number of useful purposes. When the peroxide is used as a gargle or spray in cases of sore throat, it oxidizes the diseased tissues of the throat, burning them away. It bleaches the color from many forms of organic matter. Hair is bleached to a light shade by the peroxide, and injured at the same time, while ivory, feathers, and bone all take on a whiter hue because the oxygen set free from that dioxide burns out their color.

The uses of hydrogen dioxide. — Few of the uses which have been found for hydrogen dioxide surpass its use in the restoration of old paintings. The dust and dirt of centuries of sunshine and fog dim the bright colors that the old masters used in transferring to their canvases the beauty of life and landscape. With hydrogen dioxide the modern picture-restorer can burn off the accumulations added by time, and restore life and color to the faded canvas.

Preparation of hydrogen dioxide. — Hydrogen dioxide is found sometimes in rain and air, but only to a very slight amount. It is manufactured by treating the oxide of the metal barium, barium dioxide, with either sulphuric or hydrochloric acid. It is characterized by its peculiar slightly brassy odor and metallic taste.

The uses of hydrogen in its compounds. — The part which hydrogen plays in forming acids must be considered when the latter are discussed. The many roles in which it acts in the organic world can scarcely be touched here. It is enough to say that hydrogen is found in combination in almost every one of the foods we eat as well as in the water that we drink.

It is in the fountains of oil that flow from the natural reservoirs of light and heat, and beside oil, most of the fuels contain it in greater or in less degree. Though of no wide-spread use in its elemental form, its compounds serve man under a thousand differing shapes.

CHAPTER V

ACID, BASE, AND SALT

Vinegar, ammonia, and common salt. — In every kitchen may be found three things, vinegar, ammonia, and common salt. They are three examples of three great classes of compounds that run widely through nature. Vinegar belongs with the acids. Ammonia belongs with the bases. Common salt belongs with the great general group of the salts.

Acids in nature. — Apparently in the scheme of nature there is just as much need for sour as sweet. The apple on the sunny garden wall is a mass of sour puckery juice, until the slow march of spring and summer days adds sugar to the acid of the fruit. The grape, the peach, the plum, practically the whole gamut of the fruits, are characterized in their early stages by their sour and acid taste. The lemon, the lime, and quince hold their acid throughout their existence. All contain acids which only the slow ripening of time can change. Though some of the fruits change in taste from sour to sweet from the addition of sugar as they ripen, we find yet other substances changing from sweet to sour. Milk, sweet or bland when it comes from the cow, soon changes when exposed to the air and becomes sour and acid. Sweet cider exposed to the air likewise becomes acid and turns to vinegar. Not only that, but we are constantly carrying round with us a whole laboratory full of acids in our digestive organs. When our food fails to digest we

Acids in Nature
Unripe fruit is acid. The ripened fruit is sweet because of the addition of sugar.

sometimes know it by the acid taste in our mouth. When we try to cure that type of indigestion we take bases to neutralize the acids. Hydrochloric acid, one of the great acids, is one of the liquids of the stomach, and is prescribed in its diluted form for certain diseases. Name after name could be given of acids commonly known to most of us, such as the lactic acid of milk, the citric acid of the lemon, the oxalic acid which serves as a cleanser, and the acetic acid which gives the acid taste to vinegar and therefore to the thousand pickles and salad dressings in which vinegar is used.

The inorganic acids. — Most of the acids which we know and use in the home are the organic acids which come from the plant world out of doors, but the inorganic or so-called "mineral" acids of commerce are prepared by thousands and hundreds of thousands of tons, are used in almost countless processes in shops and factory, and form one of the great staples of trade. The three important acids of this group are sulphuric (H_2SO_4), nitric (HNO_3), and hydrochloric acid (HCl). Sulphuric acid is known commercially as oil of vitriol, hydrochloric acid as muriatic acid, but nitric acid is commonly sold under the same name which designates it chemically. Less common than these acids but playing necessary parts in manufacturing come a whole range of other acids; sulphurous, hyposulphurous, tannic, tartaric, and a host of others. Each of the three acids first named is important enough to warrant a good part of a chapter to its consideration, so that the story of the making, of the properties, and of the work of the stronger acids of the world will follow later.

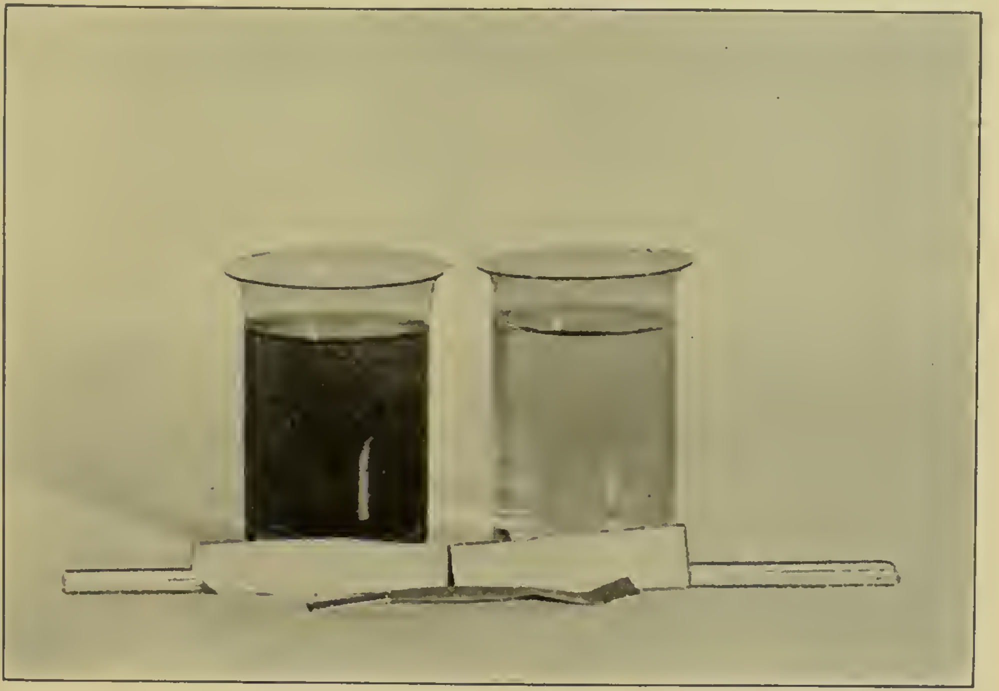
Hydrogen in the acids. — Much as the acids may vary in their individual qualities, they all possess certain characteristic qualities. They all contain hydrogen, and when this element is taken from them they all cease to be acids. The three more important acids, for instance, are called hydrogen sulphate, hydrogen chloride, and hydrogen nitrate, quite as properly as sulphuric, hydrochloric, and nitric acids. All are compounds of hydrogen with some other element or group of elements which cease to have any acid qualities when the hydrogen is taken from them. Hydrogen is the substance which seems especially to give them their acid nature, and we have already seen that the way to procure hydrogen is to displace it from an acid by means of a metal.

Non-metals in the acids. — The second necessity which every acid seems to require is that it should contain a non-metal besides hydrogen. Oxygen is commonly present in acids, and for many years was supposed wrongly to be the substance which furnished the acid quality. Oxygen is not, however, present in the important series to which hydrochloric acid belongs. Therefore we cannot class oxygen as a necessity to acid formation. We can say, however, that all acids contain some non-metal beside hydrogen. We get a second fairly common characteristic of the acids when the housewife adds baking soda and sour milk together in the making of gingerbread. The soda reduces the acidity of the lactic acid which is the acid of sour milk, and in return the soda is decomposed by the acid. Drop a bit of baking soda into sour milk or vinegar and the action shows clearly, a bubbling gas comes off and the acid is weakened. A similar action takes place when a seidlitz powder is mixed. In one of the two powders is baking soda, in another is tartaric acid. When the acid meets the soda a gas comes off and the mixture is made neutral. Not only does this action occur with baking soda, but it occurs when chalk or marble, limestone, washing soda, or potash such as comes from wood ashes, is treated with an acid. The reason for the action is this: all these substances are carbonates, substances which contain that carbon dioxide gas which (as we know from the chapter on oxygen) is formed whenever carbon is burned in air or oxygen: in general where an acid attacks a carbonate, the carbonate is decomposed, the gas carbon dioxide is given off, while some of the acid is neutralized and loses its acid qualities. One of the values of baking soda, sodium bicarbonate, is its power to neutralize some common acids.

Testing for acids. — A fruit stain on a table-cloth is the dread of a careful housewife. The acids of some fruit stains not only leave spots but affect the fabrics as well. All the strong acids such as sulphuric acid affect vegetable colors and destroy such organic substances as cloth. In fact the most valuable test by which acids may be recognized comes from their action on certain vegetable substances which is always the same. From a tiny lichen comes a liquid, the litmus solution, which furnishes us with litmus, the commonest test for this class of chemicals. Acids turn the blue of litmus to a bright red. The test is quite delicate, even a drop of a faint dilution of acid turning a bit of blue litmus paper (paper colored by litmus solution) red. Methyl orange, a yellow coal tar compound, turns red when acids are added to it. Both litmus and methyl orange belong to the so-called series of indicators, which show faithfully whenever an acid is present by some clearly defined change of color such as those mentioned above.

Defining an acid. — If we have a substance, then, which is sour, which turns blue litmus paper red, and (in the case of most acids, not all) breaks down a carbonate such as baking soda with evolution of a gas, we can be reasonably sure we have an acid substance. If we can go farther and find that it contains hydrogen and one other non-metal we can be surer still. But besides all these tests we have yet another possible characteristic, the action of acids on bases, on substances like the ammonia with which we started.

The bases. — Most of us at one time or another have had the rather unpleasant experience of being stung by a bee. Probably at the time we were unaware that we were undergoing an interesting chemical experiment, the action of formic acid on the human body. Yet that was what was really taking place. The bee's sting, the thrusting of the little pointed sword through the skin, would do little damage were it not for the fact that the bee's sting is really a tiny tube or grooved dart through which the insect injects through the punctured skin a drop of formic acid, which exerts a most irritating effect on the tissue. One of the commonest cures for a bee sting is ammonia. Why? Simply because the ammonia neutralizes the acid of the bee. Ammonia is a representative of the great class of bases which are the opposites of the acids. To accomplish similar neutralizing action the common lime-water of the druggist is given with milk to neutralize internal acid changes. All these basic substances have the property of neutralizing acids. The caustic potash which is used for soap-making is another one of these bases, and any one who has ever put his hands into a solution of caustic potash in water knows one characteristic of many of the bases, especially the strong ones, the slimy soapy feeling.



Neutralization

Hydrochloric acid and sodium hydroxide treated with the indicator phenolphthalein. The acid on the right is colorless. The base on the left is pink. Small pieces of litmus and turmeric paper rest against the beakers, together with the stirring rod, which should be used in all neutralization experiments.

Bases attack cloth and turn vegetable colors as well as acids, but, as might be expected from their action in neutralizing acids, their action is precisely opposite to that of the acid. Where acids turn blue litmus red, bases turn red litmus blue. Where acids turn yellow methyl orange, red, bases turn red methyl orange, yellow. Bases, moreover, have two indicators which belong more especially to them, turmeric and phenolphthalein. Turmeric is a solution which comes from an oriental plant like the ginger plant. Paper stained with the light yellow turmeric turns a deep red when dropped into a solution of a base. (This test is shown by one acid — free boric acid affects turmeric. This acid is seldom met in the laboratory, however, so that this indicator may be considered to belong to the alkalies.) Phenolphthalein is a coal tar compound whose clear solution turns pink when a base is added. Phenolphthalein turns colorless again when an acid is added, so like litmus it serves as an indicator for both acid and base. It cannot, however, be used in the presence of a carbonate.

How some bases are made. — Most of the bases, like lime, caustic potash, and caustic soda, contain oxygen and hydrogen in a form we have met but once before. In the study of hydrogen we found that sodium would break down water and set free some of its hydrogen, while the compound remaining behind in solution contained sodium, oxygen, and some hydrogen. That compound was sodium hydroxide. Remembering that water is made up of two atoms of hydrogen and one atom oxygen, let us look again at the reaction which occurred:

$$\text{Sodium} + \left.\begin{matrix}\text{hydrogen}\\ \\ \text{hydrogen}\end{matrix}\right\}\text{oxide} = \text{hydrogen} + \begin{matrix}\text{sodium hydrogen oxide.}\\ \text{(sodium hydroxide)}\end{matrix}$$

$$\text{Na} + \text{H}_2\text{O} = \text{Na(OH)} + \text{H}.$$

There are two atoms of hydrogen in a molecule of water. One of them in this instance went off free as a gas, and the other united with the oxygen and sodium and formed sodium hydroxide. The oxygen and hydrogen that joined with the sodium joined together very closely, so closely in fact that they formed a radical, a combination of atoms acting like a single atom which sticks together through thick and thin. This radical, which forms a hydroxide (hydroxyl it is called), is made up of one part oxygen and one part hydrogen, and plays some very important parts in the chemical drama. It is found in all the true bases. If in this case one touched the solution of sodium hydroxide left after hydrogen went off, it would be found to have the slimy feeling which characterizes the bases. It turns red litmus paper blue, and brown turmeric paper red. Sodium hydroxide ($NaOH$) and potassium hydroxide (KOH) as we have said are bases. Another striking difference between acids and bases is the almost invariable presence of a metal in a base. It is almost invariably absent in an acid.

Defining a base. — The terms base and alkali are to-day somewhat confused. Many writers use the two words practically interchangeably. The alkaline qualities of a base seem to depend upon its possession of all the hydroxyl radicals that the metal from which it is formed can hold. So we may say that a base is the hydroxide of a metal which holds no other atoms or radicals except hydroxyl. Now for one use of the term alkaline which is slightly differentiated from the term basic. We shall find that there are oxides and carbonates which neutralize acids, and that some of them have alkaline reactions.

Preparing a salt. Defining a salt. — When the right amount of strong burning hydrochloric acid is mixed with the right amount of strong burning sodium hydroxide, the base, so that there shall be no excess of either, all the properties of both acid and base disappear. There is no sour taste, no action on carbonates, no peculiar feeling to the solution left behind, no action at all on litmus, turmeric, or any other indicator. The solution remaining is neutral. The acid has neutralized the base. The base has neutralized the acid. All the distinguishing acid and basic properties have gone. Evaporate the solution thus prepared and a white powder remains: taste it carefully and it tastes strangely familiar. It is common salt, the salt that we eat on the table three times a day, three hundred and sixty-five days in the year. The action that went on might be written thus:

$$\begin{array}{ccc}\text{Sodium hydrogen oxide} & + & \text{Hydrogen chloride}\\ \text{(Sodium hydroxide)} & & \text{(Hydrochloric acid)}\\ & = & \\ \begin{array}{c}\text{Sodium chloride}\\ \text{(Common salt)}\end{array} & + & \begin{array}{c}\left.\begin{array}{l}\text{Hydrogen}\\ \text{Hydrogen}\end{array}\right\}\text{Oxide.}\\ \text{(Water)}\end{array}\end{array}$$

$$Na(OH) + HCl = NaCl + H_2O.$$

Acid + base = salt + water. — That is to say, the action of the sodium hydroxide and the hydrogen chloride has produced not only sodium chloride, common salt, but water as well, water formed from the hydrogen of the acid and the hydrogen and oxygen of the base. Such change is characteristic of hundreds of reactions and illustrates a most important general rule. When an acid is added to a base the product is a compound known as a salt (named a salt from the common salt type) and water. Acid + base = salt + water.

Here then is another characteristic of both acids and bases. When they act upon each other, water and a substance unlike either, known as a salt, is formed. We can add that to our list of acid and basic properties.

Salts. — Salts may be formed from the addition of an acid to a base or from the action of an acid on a metal or the oxide of a metal. They are as various as there are various metals to form bases, and acids to combine with them. Their characteristics vary greatly, but few of them act like acids or bases. The normal salts, those which do not contain hydrogen or hydroxyl, do not affect indicators as do these classes, nor do they possess other characteristics like them. The salt taste, something like that of common salt, is found in a number of members of this group such as epsom and rochelle salts. It may be recognized as well in the salts dissolved in such mineral water as Apollinaris and Lithia. The salts of the various acids will be considered under the headings of the metals which give them many of their differing properties. Their numbers, their lack of common characteristics, and their variation is too great to allow such general treatment as can be given to acid or base.

Normal, acid and basic salts. — There are two kinds of salt besides those so-called normal salts in which appear a metal, and a non-metal (excluding hydrogen) or the radical of an acid. These are acid and basic salts. When a metal acting upon one acid drives out only a part of the hydrogen there is left a compound known as an acid salt. When salt, for instance, acts on sulphuric acid, under some conditions a compound known as sodium hydrogen sulphate or acid sodium sulphate ($NaHSO_4$) results. This compound contains some hydrogen. Under other circumstances all the hydrogen is replaced by the metal sodium, and a normal salt sodium sulphate (Na_2SO_4) which contains no hydrogen is left. Still a third group of salts exists, those known as the basic salts. These commonly contain one or more molecules of some normal salt combined with a molecule or more than one molecule of the base. So basic lead carbonate, white lead, contains two molecules of normal lead carbonate and one of lead hydroxide. Its formula is $2PbCO_3, Pb(OH)_2$.

The uses of acid, base, and salt. — The three groups, acid, base, and salt, touch the chemical world and the world of household and factory at a thousand places. The acids are solvents, builders, and combiners. The bases are constant neutralizers and builders as well. The salts used in dyeing cover our walls with pleasant colors on hanging or paper, give brilliancy to tie or ribbon; used in food they facilitate digestive processes and build the structure of bones; used in a thousand ways they are making living easier and simpler. All in all, while their loss might not show the apparent void that the disappearance of oxygen, water, or air might leave, we could ill afford to be without any of these useful servants.

CHAPTER VI

WATER

Formula for water = H_2O

Occurrence of water. — If you recall the old-school geographies you will remember two pages where the Eastern and Western hemispheres faced each other. On those hemispheres the water of the seas was tinted a different shade than the land of the continents. What a lot of water showed on that map. Those spots of blue that showed the Pacific and Atlantic, the Mediterranean and the great lakes, those blue lines that marked rivers like the Amazon and the Mississippi, all stood for the far-reaching water which covers by far the greater part of our terrestrial globe. And when one realizes that not only the seas and the lakes, the brooks and rivers contain liquid water, but that the air holds vast quantities of water in gaseous form, and that the polar seas are vast fields of solid water, ice, we can understand that of all compounds on the surface of the earth water is the most abundant.

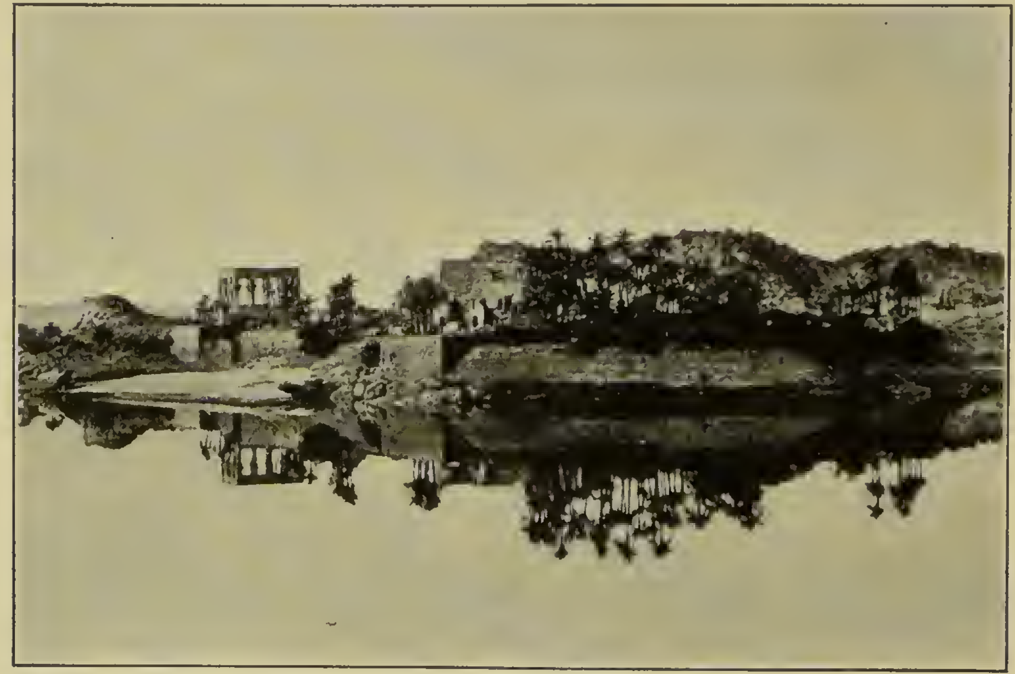
Uses of water. — Wide-spread as are the sources from which water comes, its uses are almost as various. First of all comes the dissolving power of water, its action as a solvent. Every time we put sugar into tea or chocolate we are using water as a solvent, every time we put salt into soup, or soap into water, we dissolve a solid and a solution results. The ordinary water solution is called a physical solution because the solids which are dissolved come out of solution in just the same form in which they went in, once the water is removed. Put a teaspoonful of sugar in a saucer of water. It wholly disappears. Leave the sugar solution in the warm sun for a day. The water which kept the solid spread throughout the liquid disappears, and the solid sugar remains. The liquid evaporated has changed from liquid water to gaseous water, and has passed off into the atmosphere around.

SUSPENSION

Soil brought down by a river and deposited along its course.

When any volume of water (or of any other liquid for that matter) has the maximum amount of solid in solution that it can hold in that condition it is said to be saturated. You know that you can only dissolve a certain quantity of sugar in a cup of tea. If you add more it will simply collect on the bottom of the cup. A cup of tea which has dissolved all the sugar it will hold is a saturated solution. In general, the warmer the liquid the more of a given solid it will dissolve.

Suspension and solution. — The solvent power of water, together with its carrying power, is constantly changing the surface of the earth. Great rivers like the Mississippi are constantly eating away their banks, taking the earth into suspension and bearing it miles away, finally depositing it upon the shore farther down or leaving great bars or islands in mid-stream. The Nile, overflowing its banks, floods Egypt. Withdrawing, it leaves the rich loam which it has borne down in suspension to fertilize the barren soil. The flow of brook or river, the constant dropping of the rain, the ice of the winter and the warm water of summer are gradually wearing away the soil, constantly breaking down the rocks and cliffs and leaving looser, more powdery earth to form new possibilities of fertile fields by the release of valuable plant foods which have been imprisoned.



Fertilization by the Aid of Water

The banks of the Nile which owe their fertility to the rich soil deposited upon them by the river.

The service of water to plant and animal life. — The service of water as a food to plant and animal life cannot be overestimated. Every plant gains a large part of its nourishment through the solutions which it drinks up through the tubular channels of its roots. The rain, pouring down to the earth and sinking within it, bears with it soluble plant foods from the soil. These solids in solution nourish the plant as the solids in solution in milk nourish us. We take our food through the mouth, the tubes of the alimentary canal. The plant takes a part of its food through its roots. Largely made up of water is the frame of every animal, whether it be of the human body or that of some member of brute creation. The replenishing of that water is a necessity and the source of supply is found to no small degree in fruits and vegetables which are very largely made up of water. The water that we drink washes and cleanses the interior of the body as the water of our bath cleanses the exterior, clearing both of accumulations which interfere with the proper functions of the various organs. Plenty of water inside and out is one of natures first rules for health.

Water in the foods. — Whether we drink much water or not we eat a great deal of it. It makes up far the greater part by weight of many of the solid foods we use. In most of the fish tribe it makes over three fourths of the weight that brings down the scale in the fish shop. Over sixty per cent of the weight of beef that purchasers pay for on their grocery bill is made up of the same liquid. In vegetables you are often eating ninety per cent of water. Asparagus and tomatoes, for instance, have over ninety-five per cent of water, while fruits, such as strawberries and watermelon, are over nine tenths water. Bread probably contains as little water as any of our common foods. It has about thirty-five per cent of this compound in its make-up.

Dangerous impurities of water. — Absolutely pure water does not exist in nature because of the very point mentioned a moment ago, the great powers of solution and suspension inherent in water. The clouds hold gaseous water, and the rain which falls from them might be supposed to be pure. But gases as well as solids are soluble in water, and the rain as it descends bears with it a burden of dissolved gases of the atmosphere. Tiny motes, solid particles of matter, are shown to be floating in the air by every sunlight ray. Many of these come to earth, brought down incased in rain-drops. If rain-water is not strictly pure, evidently the water on the surface of the earth stands little chance of being absolutely free from foreign bodies. Once the rain touches the earth the thousand soluble matters that cover the surface are dissolved and hurried with the water in streams to brook, pond, or river.

EROSION

A river cutting its way through rock. The cutting and grinding effect of the water is shown by the rounded rocks in the foreground.

It is by no means essential that water should be absolutely pure to make it thoroughly healthful and satisfactory for ordinary use. Only, and this a most important qualification, it must be so far pure as to carry no substance dangerous to health. Typhoid, cholera, and other contagious diseases are water-borne, that is, the germs of those diseases live and spread in water and enter the human body through drinking water. The utmost care must be taken to see that every source of water supply be safeguarded from animal contamination. Sand or earth, leaves or wood, can exercise no injurious effect on drinking water, and may even be beneficial in their addition of healthy mineral salts. It is the drainage from human habitations, from barn or stable, which, reaching drinking water, has spread pestilence again and again, and has brought death to thousands. The danger of such water-borne disease goes back once more to the carrying power of water. The rain sinks into the earth, bearing with it any germs of disease which may exist near the surface; and carries them along in the journey which rain-water always makes towards the nearest pond, lake, or river. If the water passes through a sandy soil, it becomes filtered and purified in many less feet than in a soil composed largely of clay. In the first case the harmful matters deposit quickly and evenly. In the second case the water may run for a long distance through cracks in the soil without practically any filtration. With any soil, however, the source of supply of water should be above any possible drainage. City water is now quite generally filtered in one way or another.

Soft and hard waters. — Beside the division into clean water and that containing harmful impurities, waters are divided into soft and hard, according to the substances which they carry and especially as regards their action on soap. In soft waters, soap readily lathers and exerts a rapid cleansing action. In hard waters, soap lathers only with difficulty. Often it will not lather satisfactorily at all. In some cases, if the water is boiled the hardness disappears and the soap acts as in soft water. In yet other cases boiling has no effect. Such a condition is wholly due to the action of the solids upon the soap. Hard water contains either magnesium or calcium, that is, it holds compounds of the metals which are found in magnesia and in limestone.

The "brown fur" of a kettle. — In one place we can commonly see the solids that hard water holds: where hard water has been boiled for some time in a kettle a brown "fur," as it is called, will appear on the bottom and sides. This is really a crust formed when the solids which the water held came out of solution. The "fur" is chiefly made up of calcium carbonate, which is the same substance that appears in limestone, in chalk and marble. It generally contains beside this substance some magnesium carbonate or magnesite, which occurs in soil in somewhat less quantity than calcium carbonate. Both these compounds are white, but they are tinged brown by iron rust and by the vegetable matter of the water.

Permanent and temporary hardness. — If carbonates only are present, water, when boiled, is likely to become soft, for water containing carbonates generally contains carbon dioxide, and loses most of the gas when boiled. Some of these hardening compounds, soluble in water containing carbon dioxide, are insoluble in pure water. When the kettle boils the water loses its carbon dioxide, the solid carbonates sink to the bottom and the water above is left soft. When hard water is made soft in this way by boiling it is said to have temporary hardness. When compounds of lime and magnesia, such as sulphates, are in the water they will not deposit in this way on boiling. Water in which such compounds appear is called permanently hard, since it cannot be softened in any simple way.

Soap in hard water. — So much for the burden which hard water carries. Now for its effect on soap. In order to cleanse, soap must decompose, must separate into some of its constituent parts when placed in water. Lime and magnesia compounds form new insoluble substances with soap, and thus prevent its separation. The use of hard water therefore means a constant loss of soap in washing. Before the soap will lather, before it will clean, it must combine with the solids present in the water.

The danger of using hard water in a boiler. — While some hard water can be made soft for the wash-tub by boiling, and the solids deposited can be cleansed from a kettle, hard water in a steam boiler is a constant trial. Layer after layer of tiny crystals falls on the bottom and sides of the boiler as the water within boils into steam. Here and there in spots and seams heavy loads of the powdery compound accumulate. The iron of the shell conducts heat readily, but the crust formed conducts it badly, and, in its vain endeavor to pass heat through, the iron of the boiler becomes hotter and hotter. The danger of this can be seen when one realizes that the heat of the fire ordinarily passes through the iron into the water of the boiler, and its energy is used in forming steam. When severe incrustation from hard water takes place the heat is retained in the iron, passes through the crust only with difficulty and easily brings the boiler to a red heat. Should the crust crack under that heat and allow the water within the boiler to reach the heated shell, a sudden burst of steam results and the boiler explodes. In consequence great care should be exercised in freeing boilers from this crust where hard waters are used.

Stalactites and stalagmites. — Water trickling through beds of limestone produces curious grottoes. Wonderful caves whose dark recesses show cone and pillar, arch and grotto, come from water-bearing calcium carbonate which leaves its burden in the cavern. The cones of limestone which hang downward from the cavern roof are called stalactites. Those which reach upward to meet them rising from the ground are called stalagmites. Stalactites and stalagmites may both be colored with beautiful reds, soft grays and yellows by iron and vegetable matters which are borne in with the water.

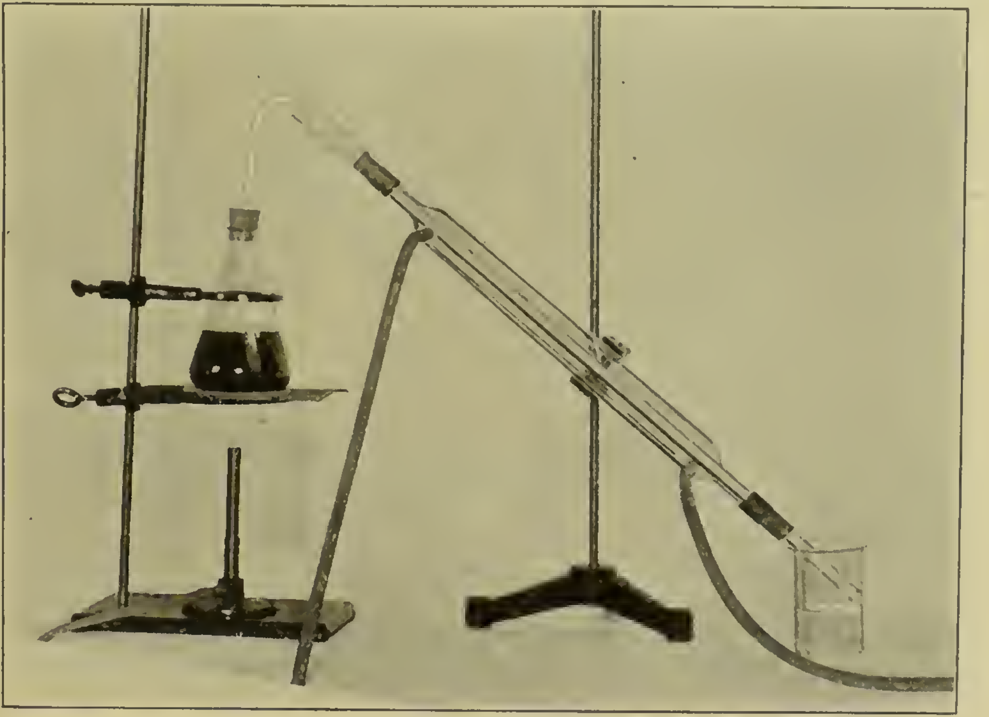
Water of crystallization. Efflorescence. — Not only does the water carry substances in solution, taking them up and laying them down, but it also becomes an integral part of many substances in the form known as water of crystallization. Take ordinary dry powdered washing soda and leave it in saturated solution in water. White crystals form which weigh considerably more than the dry powder. Water of crystallization has entered into them and has become a part of them. Leave the crystals formed from the solution exposed to the air for a while and they soon begin to lose the water and crumble back into dry powder. Crystals, which lose their water in this way at the ordinary temperature of the air, are said to effloresce. On the contrary, when dry substances attract water and take it up from the water vapor of the air they are said to deliquesce: substances which exhibit an attraction for water are called hygroscopic. Gases or other substances from which water is removed are said to be dehydrated.

Deliquescence. — Potash and chloride of lime both belong to the deliquescent group of substances. Expose either of them as a dry powder to the air and they soon become moist and sticky. They have taken up water from the air and have deliquesced.

Natural lithia waters. — The lithia of lithia water, like the various salts of mineral and spring waters, comes from solvent water trickling into mineral springs through deposits of lithia or other solids and carrying them on in solution. The effervescent spring waters in which bubbles of gas rise through the liquid are charged naturally or artificially with carbon dioxide.

Distillation. — To free water from the solids which it carries and to procure practically pure water in the laboratory is by no means difficult. If we boil water the liquid turns to steam or gaseous water, and any solids within are left behind. If we then cool the steam and condense it back to liquid water, we shall have water free from impurities. This is the principle of the common process of distillation in which water is turned into steam and steam is turned back into water. By this process the salt water of the ocean may be made to yield as pure water as the reservoir on the land, and that with very simple apparatus. Something in which to boil the water, a tube through which the steam may pass, and some means of cooling that tube so that the steam may be condensed, is all that is necessary.

Proving the composition of water by analysis. — The fact that water is made up of hydrogen and oxygen in the proportion of two parts hydrogen to one of oxygen can be readily proved by either analysis or synthesis, by either breaking it down or building it up. To do either one, that valuable agent of the chemist, the electric current, is used. Send an electric current through water containing a few drops of sulphuric acid to make it a good conductor and the water rapidly breaks down into its component parts, hydrogen and oxygen. The current enters and leaves the liquid on poles of the metal platinum. If tubes are placed above the place where the current enters the water, hydrogen will rise in one tube, oxygen in the other. Twice as great a volume of hydrogen will rise as of oxygen, showing by analysis that water is made up of twice as much hydrogen by volume as of oxygen.



DISTILLATION

The liquid in the flask, turning to gas, passes down the inner tube of the condenser (Liebig Condenser), is turned to liquid by cooling water flowing through the outer tube or jacket, and falls as liquid into the beaker at the end. The cooling water enters by the lower rubber tube and passes out by the upper rubber tube. Any impurities present in the original liquid are left behind in the flask.

Proving the composition of water by synthesis. — If we reverse this process and, having filled a tube called a eudiometer with three parts hydrogen to one part oxygen, explode the mixture by means of an electric spark, we shall find that just two parts only of hydrogen have united with the oxygen, while the remaining hydrogen is left undisturbed. So by synthesis we can show that the composition of water is two parts of hydrogen to one of oxygen. $O_2 + 2H_2 = 2H_2O$. This experiment is not safe for ordinary laboratory performance. It can only be done with care on the lecture table.

The expanding power of gaseous water, steam. — On one of the properties of water depends to a large degree the thousand and one luxuries and necessities of our daily life, the power which steam possesses of expanding and doing work when liquid water is turned into gaseous water or steam. The steam-engine depends first, as we shall see, on the oxidation of carbon, on the fact that when carbon unites with oxygen a large amount of heat is given off. Secondly, it depends on the heat generated turning the water of the boiler into steam. Third, it depends upon the expansion power of that steam or on the capacity which it possesses of moving a piston or a set of pistons to and fro. From the power that is generated by that movement of pistons back and fro comes the livelihood of most of the indoor workers of the world, comes the turning of the railroad wheel as it speeds across the continent, comes the mighty onthrust of the battleship's turbines as she rushes across the deep.

The unusual expansion of water. — In the change from liquid water to solid water, ice, we find properties quite as important as the change from water to steam, much as that means. Most substances expand when heated, contract when chilled. Water, on the other hand, follows this general principle to a certain point and then does just the reverse. It contracts as it cools until it reaches a point about 4 degrees Centigrade or about 40 degrees Fahrenheit (that is to 8 degrees Fahrenheit above the freezing-point, which is 32 degrees Fahrenheit). Then it begins to expand. It expands regularly down to the freezing-point. There it turns into ice. The lightness of ice, the fact that when water freezes it expands, is evident every time you see a glass of ice water. The ice floats well above the water much as wood floats. Its density, then, must be less than that of water, and that lowered density must have come from the expansion of the ice.

THE GLACIER OF THE RHONE

A frozen river of solid water. Here ice, because of its expansion, occupies far more space than liquid water would require.

The effect of water's expansion on our existence. — On the fact of the expansion of water from 40 degrees to 32 degrees, and on the fact that it expands still farther as it changes to ice, depends our very existence. Were it not for this a huge layer of ice and snow would cover fertile field and busy town, lake steamer and railroad train, forming a polar cap which would reach down covering the temperate zone and all but a narrow strip about the equator. Think what a scene of desolation that would be! An Arctic winter with its stillness would hold chained in an icy grasp all our busy world. All this would be were it not that water goes contrary to the common experience of substances, expands where it would be expected to contract, and contracts where it would be expected to expand. Let us try to unriddle this puzzle.

How water expands in a lake. — A lake on a late fall day may show a quiet and placid surface apparently without the movement of a ripple, yet it may be in constant motion. As the temperature falls towards the freezing-point the cold acts first on the surface water which contracts steadily in the beginning and therefore grows heavier than the water beneath the surface. A liquid of greater density will, as we know, sink in a liquid of lesser density as mercury sinks in water. When the surface water contracts its volume becomes less while its weight does not change. In consequence its density becomes greater and it sinks. Once this water has sunk the lighter water from the bottom rises to take its place. This cools and sinks in its turn. Such changes, such rising and falling of liquids of different densities, are known as convection. In this way the temperature of the whole lake grows steadily colder until about 40 degrees Fahrenheit. Then comes a sudden change. As the temperature drops towards the freezing-point, the surface water expands instead of contracting, becomes lighter than the water underneath and, being of less density, floats like oil on water. That surface layer remains at the top till the freezing-point 32 degrees Fahrenheit is reached. Then at some blast of northern wind the floating liquid becomes chained into ice, expands and bars the cold from the warmer water beneath. All winter the lower water shielded from the cold never reaches the freezing-point, and the fisherman cutting a hole through the ice finds, not far below the surface, black water whose temperature is well above the freezing-point.

The effect of a sudden change in the expansion of water. — Suppose this irregular expansion (called, by the way, the anomalous expansion of water) did not hold. Suppose that water contracted steadily until it reached the freezing-point. All the water of our lakes would reach the freezing-point together, and would freeze from top to bottom into one solid cake of ice. Suppose you could magnify such a cake as the ice man brings by as many million times as Lake Superior is bigger than a fifty-pound piece of ice. That would make the ice cake of one lake alone. Brook and river, pond and lake, would show only gleaming lines and spots of silvery ice. The water of the city reservoirs would be congealed, and the only water we could use would be melted from ice. Through the whole long summer the sun at its hottest could melt only a comparatively few feet of the deep lakes, leaving a solid depth of ice reaching to the bottom, a reservoir of cold which would soon turn the temperate zone into an arctic wilderness.

All the possibilities of civilized life then, everything which differentiates the people of the temperate zone from the Esquimaux, comes from the simple fact that water, expanding through a few degrees of low temperature before freezing, leaves the great mass of water liquid.

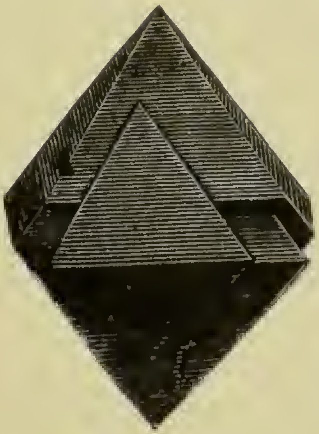
Crystallization. — One more word concerning water's characteristics: The snowy crystals of the window-pane, the forests and sprays of ice that coat the glass on a frosty morning, are the most usual crystals we have. They show the common characteristics of crystals in their regularity of form and shape, their smooth even surfaces meeting in straight lines. Many bodies beside water have this power of crystallization, of forming crystals, especially when they come out of a solution which has been slowly evaporated. Under such circumstances, the solid commonly, but not always, unites with water of crystallization to form the regular crystal. Rock crystal is one of the best known of the natural crystals, and one which gave the first impetus to the study of crystallography, as the study of crystals is named. Many substances are freed from impurities by crystallization, as under such circumstances impurities will not form in the crystals, but are left behind in the bath (called the mother liquor) from which the crystals come. Since some substances crystallize at higher temperatures than others, two compounds may be separated from the same solution by means of successive crystallizations.

Ice Formation

The crystallization of water. The commonest crystals that we meet.

The uses of water. — From the water we use for cleansing the body inside and out, from the thousand manufacturing processes where this liquid is necessary to development, from the surface of the earth changed and furrowed by the ardent strenuous labor of water, ice, and steam, from the busy engines and water wheels that hasten the march of trade, from the water of tributary and reservoir, of ocean and of lake, we can get finally some conception of what this great servant of ours does for the common weal. Ever laboring, ever changing, ever upbuilding, few indeed are the forces which compare with the water which we throw aside so often as a thing of no account.

A Crystal



Another Crystalline Formation

CHAPTER VII

LAWS

Laws made by man. — To most of us the word Law stands for very definite realities. It means the broad, white Capitol at Washington, or the long, low pile of the Houses of Parliament on the banks of the Thames. It means the statutes drawn up in those places by the legislators within, or else it means the municipal courts and the policemen, which enforce those statutes. All laws of this type are made by man. Some of them are obeyed well, some ill. The whole machinery of the law is called into service to protect the property of the citizen from attack, and yet day after day we read of burglaries. No attention at all, on the other hand, is paid to some of the laws which stand forgotten and unrepealed for centuries on the pages of the statute books.

The laws of nature. — In so far as the enforcement of these man-made laws depends on men, they are fallible and weak. The laws of nature, the great laws which underlie natural science on the other hand, are infallible, unerring, so that whether we wish or no they require obedience. For those who try to disobey them, the punishment is certain, — injury or death results. High up on the snow of the Alps stand shelves of snow whereon crossing may sometimes be made. If the ice holds strong, the mountaineer can cross. If the sun melts and weakens the shelf of snow so that it can no longer bear the weight of the climber, the man who tries that path plunges headlong to destruction. He has disobeyed the laws of nature. The force downward of gravitation was greater than the supporting force which acted upward. And whether the climber took the path recklessly or ignorantly the result was the same. He paid for his ignorance or daring with his life. Nature never excuses ignorance any more than recklessness. The laws of nature under which we live can never be disobeyed with impunity.

The demonstration of nature's laws. — Before these great underlying laws of the universe such as that of gravitation (the law which all things in the universe obey, by which the apple falls to the ground and the earth moves in its path about the sun) could be stated in definite terms by man, experiment after experiment was necessary, year after year of observation. Caesar, when he pushed his legions into Gaul to attack the German tribes, was able to progress more rapidly when the leaves had fallen from the trees of the great forests. Those leaves that fell two thousand years ago dropped because their life was done and because new life had come. This autumn in the forests where Caesar's legions marched the leaves fell just as they did when they fluttered down upon the Roman Eagles passing underneath, just as they have fallen in each of the two thousand autumns that have passed between. It is an unchangeable law, a law proved by all of human experience, that the life of the leaf shall perish after its allotted time is run and that new life shall come. We cannot escape these laws which rest upon such foundation. We can learn them and put them to our own use. In ignorance of their action lies danger. In knowledge of their action lies power, and a further recognition of the wonder and beauty of the universe.

The law of the conservation of energy. — The law of the conservation of energy, which states broadly that energy can never be destroyed, is one of the underlying laws which the world always obeys. Energy is, perhaps, even more difficult than matter to explain briefly and simply. It has been defined as the capacity for performing work. It is shown by some action in which matter moves or undergoes some change. The flatiron, growing warm upon the stove, is receiving heat energy. The incandescent filament of the electric light gives out light energy produced by the electrical energy which heated the wire. A hammer falling on a nail is an example of mechanical energy. A vibrating piano string sends forth the energy of sound. Chemical energy produces heat energy when carbon oxidizes to carbon dioxide. But whatever type of energy we have this law holds good. *In any separate action,* just so much energy as you put in, just so much energy you will take out. When we put a pound of coal into a furnace or a stove, it possesses, by its power of combining with oxygen to form carbon dioxide, a certain quantity of heat which lies latent, hidden till flame turns the black solid to colorless gas. Just that amount of energy, no more and no less, will serve to warm the stove, heat the room, or cook the dinner when the coal is burnt. Not all nor anywhere near all that energy will be available for cooking the dinner, — the equation will always read:

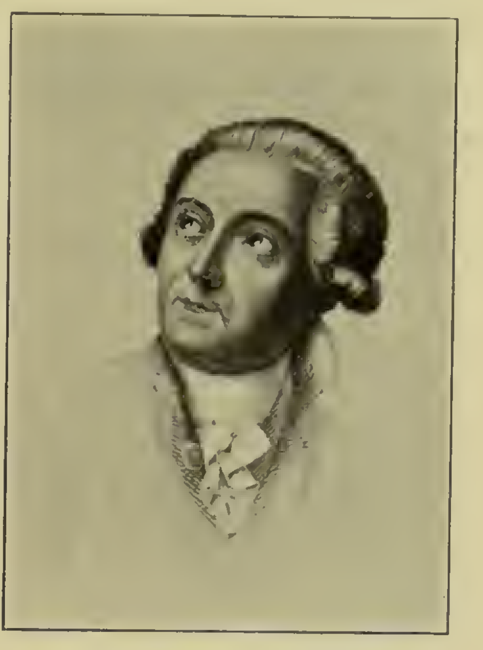
$$\text{Total energy} = \text{useful energy} + \text{waste energy};$$

but this second equation always holds good,

The amount of energy you put in = the amount of energy you take out.

Applications of the law of the conservation of energy. — In every action, be it physical or chemical, this law of the conservation of energy persists, and in it can be found the reason why perpetual motion is impossible. No machine will give out more energy than is put into it. When its supply of energy is exhausted the machine stops. This law is true in the action of the human mechanism as elsewhere. Our food is our fuel. If we take the proper food and eat it properly it provides us with energy to do our work and play. Brain energy is given out for study, muscular energy for play. We can get no more energy out than we put in. A one-sided diet, like a diet of tea and toast, for example, will not supply the energy which this body of ours requires. It is absolutely essential that we should have a well-balanced, sufficient diet, if we are to do the work which we are meant to do in the world.

The law of the conservation of matter. — Next to the law of the conservation of energy comes the second great law, the law of the conservation of matter. Matter is indestructible, says this law. Just so much matter as goes into any separate action, just so much matter comes out of it. If we put twelve pounds of carbon and thirty-two of oxygen into any reaction, we get forty-four pounds of carbon dioxide gas out of it. Carry this law up to its largest viewpoint, and you will see that we are forced to believe that the sum total of the matter of the universe is a constant quantity, that it can never change, that there is now, has been, and always will be the same amount of matter making up the whole. Its form may change. The oxygen of the air which blew into your open window last night may have gone into a glowing furnace, turned to carbon dioxide, gone from that form into the velvet petal of a rose, and from there through various transformations, ending next week in combination as a bit of iron rust on a roof top. It may change in a year into a hundred thousand forms, but it never can be destroyed. No matter what form it holds it never disappears. We must believe that the same oxygen which cooled the weary slaves building the pyramids of Egypt thousands of years ago is to-day somewhere on the earth either in its own form or in that of one of its varied compounds. Matter is indestructible. It can never perish.



Lavoisier. (1743–1794)
The Father of Modern Chemistry. A great investigator.

The law of definite proportions by weight. — Closely bound up with these marvelous general conceptions are the laws on which the foundations of chemistry stand. The law of definite proportions by weight says, "The same compound always contains the same elements combined together in the same proportions by weight." Let us see just what that means. Take a bucketful of water out of the sea where the tide washes the huge piles below the tall buildings of New York. Take another where the Mersey runs into the channel near the great docks of Liverpool, and still a third where the warm waters of the East wash the black supports of the old wharves at Singapore. Now take these three buckets, taken from three corners of the globe, and boil away the water. White salt will be left behind, sodium chloride. It is made up of two elements which we shall consider shortly, the metal sodium and the non-metal chlorine. Try by every test known to man to find a difference in the three samples of salt taken thousands of miles apart and you will be unsuccessful. Each one will contain exactly the same amount of sodium and exactly the same amount of chlorine put together in the same proportion by weight to make up the salt. Not only in the salt of the sea, but in the salt of the rock salt mine, in the salt of the saline spring, and in the salt which we make in the laboratory, we find the same thing is true; we invariably get just the same proportions of sodium and chlorine to make sodium chloride, and no matter how hard we try to make a greater or less amount combine, we *never* can do it. The sand of the beach and of the soil is just as invariable in constitution, whether we find it on the beach of Iceland or of Brazil, in the depths of the earth or on the tops of the mountains. Every known compound obeys this law, every combination of oxygen, nitrogen, and chlorine, every union of the elements. All combine in definite (*or constant*) proportions by weight.

The law of multiple proportions by weight. — Grouped with the law of constant proportions by weight we find the law of multiple proportions by weight, a law which states, "When the same two elements combine together to form more than one compound, the different weights of one of the elements, which unite with a constant weight of the other bear a simple ratio to one another," or stating it differently, "When one element unites with another in two or more different proportions by weight, these proportions are simple multiples of a common factor." Here is a perfectly simple fact expressed in somewhat cloudy terms. Suppose we have a stable where there are always kept two cows. Sometimes there is one horse in the stable, sometimes two, sometimes three, sometimes four, but whichever there is, the number of horses is always a simple multiple of two, the number of cows, 2 × 1, 2 × 2, 2 × 4. We can never have a horse and a half or two horses and a third. We must always have whole animals, and the number of horses always bears a simple integral relation to the cows. They are always simple multiples of the cows. The number of one kind of animal is always in a simple ratio with the other. We never find a fractional ratio such as 1.6 to 1 or 2.4 to 2. Whole numbers invariably appear.

Illustrations of the law just cited. — Four simple examples of this truth in chemistry instantly spring to mind. Water, the oxide of hydrogen, has two atoms of hydrogen to one of oxygen. Comparing their atomic weights we find that they unite in the ratio of 2: 16. Hydrogen dioxide, the other oxide of hydrogen, has two atoms of hydrogen in atomic weights, to two of oxygen; 2: 32. The different weights of the oxygen which combine with a constant weight of the hydrogen are in the simple ratio of one to two; 16 to 32. The same fact holds good in the two oxides of carbon. Carbon monoxide contains one atom of oxygen; 12: 16. Carbon dioxide with exactly the same quantity of carbon has two, 12: 32. Again the oxygen combines in the simple ratio of one to two, 16 to 32. When we compare the relative weights of the five combinations of the two elements, nitrogen and oxygen, we find that every time the oxygen combines in a higher simple ratio.

Here are the five oxides of nitrogen and the way they combine.

In Nitrous oxide	Nitrogen: oxygen: : 28: 16
In Nitric oxide	Nitrogen: oxygen: : 28: 32

In Nitrogen trioxide	Nitrogen: oxygen: : 28: 48
In Nitrogen peroxide	Nitrogen: oxygen: : 28: 64
In Nitrogen pentoxide	Nitrogen: oxygen: : 28: 80

That is, if we take twenty-eight parts of nitrogen as a fixed quantity in the beginning, we shall find that the oxygen added to it is as 16–32–48–64–80. Now divide all those figures by 16, the first weight of the oxygen added.

$$\tfrac{16}{16},\ 1,\ \tfrac{32}{16},\ 2,\ \tfrac{48}{16},\ 3,\ \tfrac{64}{16},\ 4,\ \tfrac{80}{16},\ 5,$$

There is a simple ratio of the element oxygen which combines with a fixed quantity of the element nitrogen in the simple multiples of one, two, three, four, five. No matter what weight of the oxides of nitrogen we may take, the oxygen is always present in them in this simple ratio. If we put it into letters and made $2a$ the constant quantity and b the varying quantity, we should have $2a\colon b–2a : 2b–2a : 3b–2a : 4b–2a : 5b$. The b's would be simple multiples of each other.

The law of reciprocal proportions. — The third general law of chemistry (the Law of Constant Proportions being the first one mentioned, the Law of Multiple Proportions, the second) is called the Law of Reciprocal Proportions, or Law of Equivalent Proportions. It states that "the weights of different elements which combine separately with one and the same weight of another element, are either the same as, or are simple multiples of, the weights of these different elements which combine with each other," or in other words, "the relative proportions by weight in which the elements A B C D, etc., combine with the constant weight of another element X, are the same for combination with any other element Y."

Illustrations of the law just cited. — Let us see how this law fits in with an illustration of horses in the barn. Suppose we have a pair of white horses in one barn and a pair of bays in another. If we want to make a four-in-hand we can combine the two pairs and make them two and two, or we can bring up a single bay leader from yet another barn and make five, or else another white pair and make six horses in all. But we shall never have any horses and a half, or horses and three quarters. The horses will always combine to make up the teams by simple combination, by the addition of some simple multiple of the single horse which is the unit.

Sulphur and oxygen both combine with hydrogen forming hydrogen sulphide, hydrogen peroxide, and hydrogen oxide. In both the former compounds the weight of oxygen and sulphur which combine with a fixed quantity of hydrogen is the same. We have thirty-two grams of oxygen to thirty-two of sulphur, 32 : 32 : : 1 : 1. In hydrogen oxide sixteen parts of oxygen combine with the same amount of hydrogen as we find thirty-two parts combining with in hydrogen peroxide. Therefore the oxygen combines with the hydrogen in simple multiples, in the ratio of 1: 2. In hydrogen sulphide we have twice the weight of sulphur, 32, that we have of oxygen in hydrogen oxide, 16. There is another case of a simple multiple 2: 1: : 32: 16. When sulphur and oxygen combine together in sulphur dioxide, they combine in equal weights. Thirty-two grams of sulphur combine with thirty-two grams of oxygen. Their combination with each other gives the same weights which they had when they both combined with hydrogen. The law of reciprocal proportions holds good.

When sulphur and oxygen combine with the metal sodium, thirty-two parts of sulphur unite with forty-six parts of sodium to form sodium sulphide, while sixteen and thirty-two parts respectively of oxygen unite with forty-six parts of sodium to form sodium oxide and sodium peroxide. The relative proportions in which sulphur and oxygen combine with a constant weight of hydrogen, 32, 32, and 16, are the same with which they unite with a constant weight of sodium. The law of equivalent proportions holds good here and elsewhere.

The laws of chemical composition and the atomic theory. — The relation of these three great laws to the atomic theory is very close. As expounded by Dalton, the English philosopher chemist, that theory states that all chemical combination takes place between atoms which act as if they were indivisible. Each atom it is believed has a relative fixed weight which is obtained by ascertaining the amount by which it is heavier than the hydrogen atom. Without attempting to find a means of measuring the weight of the hydrogen atom we can give that weight a name and call the weight of one atom of hydrogen 1 microcrith. The weight of an atom of oxygen is sixteen times that of an atom of hydrogen, so 16 microcriths is taken as the atomic weight of oxygen. Every elementary atom in like fashion has its atomic weight based on the amount which it weighs in comparison with the weight of the hydrogen atom taken as a unit of 1. Assuming that atoms are indivisible, and each has a fixed unvarying weight, let us see the resulting action as regards the three laws. The conception that atoms are indivisible agrees with the law of constant proportions, for every compound must invariably be made up of the same elements combined in the same proportions by weight. Every given compound, by the atomic theory, is made up of the same atoms which are exactly the same wherever found, and have always the same weight. In like manner in the law of multiple proportions the atoms cannot combine in any ratio other than a simple one, for the only way you can add a unit is to add one at a time as bricks are added. When you have one atom of oxygen and wish to increase you can only do it by adding one atom of oxygen weighing sixteen microcriths more. You cannot add less. You can add more, but only single atoms or multiples of single atoms. Hydrogen oxide

John Dalton. (1766–1844)

More than any other man the founder of the atomic hypothesis.

has two parts hydrogen to sixteen of oxygen. Hydrogen peroxide has two parts hydrogen to thirty-two of oxygen. It is as if you had pound weights on a scale, you add one and the bar shows sixteen ounces, you add another and the bar shows sixteen ounces more. The conception that atoms act as if indivisible, and that each atom has a fixed weight, is supported by this law.

When we consider the law of reciprocal proportions it seems evident that if we take an atom weighing thirty-two microcriths from one substance to add it to another, the atom can neither gain nor lose in the transfer; that if two elementary atoms weigh sixteen and thirty-two respectively, in separate compounds, they must weigh sixteen and thirty-two together, or else they must form some simple multiple of sixteen or thirty-two. The atom can only be multiplied by two, three, or some simple number, for it only combines by adding units or atoms, one by one, as a bricklayer puts one brick on another in forming a wall.

The relation of the great chemical laws to every-day life. — The great unvarying laws support the atomic hypothesis and stand at the foundation of modern chemistry. Stern and cold in themselves, the possibilities they have brought to commerce by the reduction of manufacturing processes under the laws of exact science, instead of leaving them to the caprice of chance experiments, can scarcely be overestimated. Were it not for such laws, we should still be throwing heaps of various substances into vats and furnaces, spoiling as much in value as we gained, and wondering each time we carried out any manufacturing process whether or not we should get a satisfactory product. With the aid of these laws we can know that exact quantities will give us exact results. We can expect success every working day in the year.

CHAPTER VIII

CHLORINE

Symbol of Chlorine, Cl

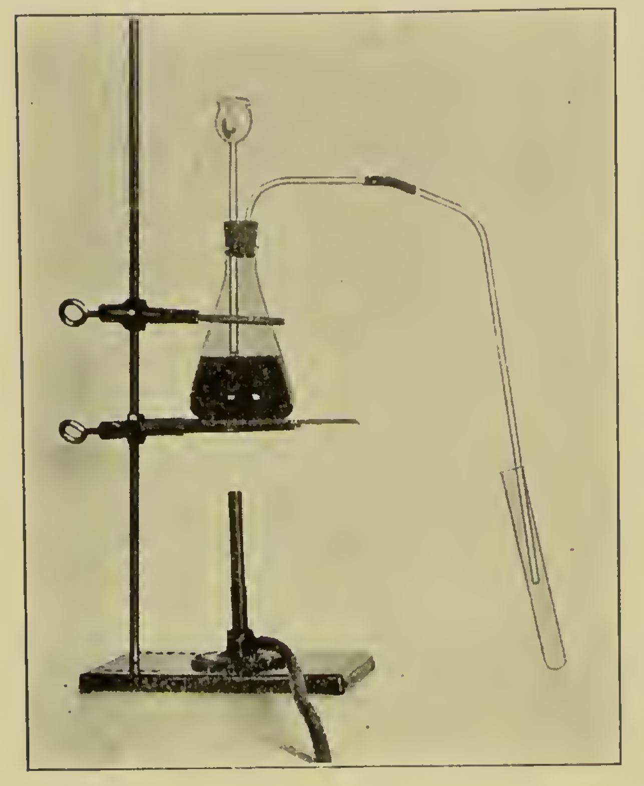
Chlorine the bleacher. — Many a tint of dress or ribbon comes from our busy looms to-day. Pale rose and deepest blue, soft pink and gorgeous crimson, shades by the thousand such as emperors never dreamed of a hundred years ago, fill the wide shop windows. Streets lined with stores that vend fresh linen, shimmering silk, or lustrous velvet are carnivals of color. Could all their colors be made permanent how great a gain would have been secured, but in how many cases do they swiftly fade? Permanence of color and laundry work seem seldom to go together. Wash goods go to the laundry in full possession of their color and dye. They come back whitened, weakened, and bleached. The agent which indirectly rots the cloth and removes its color is the gas which bleaches, chlorine.

Occurrence of chlorine. — Found nowhere free and uncombined in nature, chlorine is met the world over in the form of its compounds, the chlorides. The salt of the sea, the muriatic acid of commerce, the bleaching powder of laundry and mill, the potassium chloride of the salt mines, all contain large quantities of chlorine. Many of these compounds give up their chlorine with considerable ease, but none of them part with it more readily than does hydrochloric acid, made from the chlorine of common salt, sodium chloride ($NaCl$).

Preparing chlorine by the electric current. — When sodium chloride and sulphuric acid are heated together a strong acid compound of hydrogen and chlorine, hydrochloric or muriatic acid (HCl), results. We shall consider this compound in detail in the last part of this chapter. For our present purposes all we need to know about it is that it is produced from salt, is composed of hydrogen and chlorine, and furnishes chlorine more easily than any other substance. If we pass that willing worker the electric current, on carbon poles, through a solution of hydrochloric acid gas, exactly as we passed it through water on platinum poles, striking phenomena appear. Bubbles of gas rise swiftly from both poles where the electric current meets the liquid. In one tube rises a gas, which, tested by a burning taper, gives a blue flame and generally a slight report (though perfectly pure hydrogen gives no report). That gas is hydrogen. In the other tube rises an equal quantity of greenish yellow gas of striking appearance, chlorine. For the first time we have a gas that we can see. It is not safe to smell of the gas directly, but if a slight quantity escapes into the room we get an odor like that detected in bleaching powder, a strangling, intensely disagreeable odor which rouses a strong desire to cough. This gas is far heavier than any we have met, being two and a half times as heavy as air. It is moderately soluble, and when collected over water will not come out into the receiver until it has saturated the water through which it passes. In consequence of its solubility we generally collect it by downward displacement.

Preparing chlorine from hydrochloric acid and manganese dioxide. — The preparation of chlorine by the electric current, as explained above, shows the gas and its properties clearly, but is not so available for general laboratory use as is the preparation from hydrochloric acid and manganese dioxide (MnO_2). Hydrogen, wherever possible, tends to unite with oxygen to form water. Manganese dioxide contains considerable oxygen which it gives up fairly readily. When we heat the two substances together in a flask (*under a hood or closet, from which all fumes go up a chimney*) the manganese dioxide gives up its oxygen to the hydrogen to form water, and, taking some chlorine into combination, sets free the rest which goes off in the form of gas, which can be collected by downward dry displacement. We may express this as an equation.

Manganese dioxide + hydrogen chloride = hydrogen oxide (water) + manganese chloride + chlorine (gas). In this equation the two oxygen atoms of the manganese dioxide form two molecules of water (hydrogen oxide) with the four hydrogen atoms of the four molecules of hydrochloric acid. The four chlorine atoms of the acid separate into two parts. Two join with the manganese to form manganese chloride, and two go off as gaseous chlorine.



Preparation of Chlorine

An Erlenmeyer flask containing hydrochloric acid and manganese dioxide is placed on asbestos, as shown above, and heated gently. The gas is collected by downward displacement in the test tube. Hydrochloric acid is prepared in a similar apparatus by heating sulphuric acid and salt in the flask. Both these gases should be kept *under a hood.*

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2$$

Instead of using hydrochloric acid we can use salt and sulphuric acid with manganese dioxide. The action is the same, save that the hydrochloric acid is made simultaneously in the generator with the chlorine, instead of being made outside. In either case the fumes of the chlorine must not be allowed to escape into the room.

Testing chlorine. — Once we have the greenish yellow gas in the receiver, the next step is to test it. A flame goes out when there is nothing but chlorine present, but if some air is allowed to enter the receiver, the burning splinter may burn dully. Dry litmus paper changes not at all in the gas. Wet the litmus and both its red and blue change to white. Thrust into a bottle of chlorine a dry piece of cloth and no effect is seen. Thrust in the same piece wet, and its tint swiftly fades, the amount of color lost depending on the fastness of the color and on the goods which are dyed. Evidently, then, chlorine will bleach wet colored substances, but will not bleach dry articles. There must be some basic principle in this.

The tendency of chlorine to unite with hydrogen. — The principle is found in the extreme inclination of chlorine to unite with hydrogen wherever found. If chlorine and hydrogen are mixed in the sunlight they explode violently, forming hydrochloric acid gas. Steam is broken up so rapidly by chlorine that oxygen comes swiftly off in large quantities, leaving hydrochloric acid gas behind. Like steam, many another compound which contains hydrogen, ammonia gas, for example, is broken down in presence of chlorine, its hydrogen uniting with chlorine to form hydrochloric acid gas, its other constituents passing off.

Bleaching. — Bleaching is simply a carrying out of the general principle that chlorine will unite with hydrogen whenever it gets a chance. The wet litmus or cloth is bleached by chlorine because it is covered with water, hydrogen oxide. The chlorine meeting the water seizes instantly upon its hydrogen, and sets oxygen free. The free oxygen oxidizes the coloring matters of the cloth and burns them up. Chlorine bleaches indirectly. It takes the color from the cloth, not because it combines directly with the dye, but because it sets free oxygen in which the coloring matters may oxidize or burn. To apparently cleanse fabrics with chlorine is far easier and quicker than to wash them carefully with soap, and, as a result, many laundries use washing powders whose cleansing properties depend largely on the chlorine which they carry. Such laundries increase the business of the cloth merchants, for these powders not only bleach out the colors but burn the fabrics as they bleach, and so rot the cloths.

Chlorine in solution. — When chlorine water, a saturated solution of chlorine in water, is taken from the dark into the light, chlorine's attraction for hydrogen is seen in yet another case. In the dark this solution remains neutral, has no effect on litmus, and no acid reaction. It does, however, bleach and it exhibits most of the usual properties of chlorine. Place the solution in the light and the chlorine soon begins to join with the hydrogen of the water, while bubbles of oxygen rise to the top of the jar. Test the solution after an hour, and it will turn blue litmus red. It is acid. Try the tests for chlorine on it. It will no longer bleach. Try the tests for hydrochloric acid on it and it responds instantly. Hydrochloric acid has been formed. The chlorine has united with the hydrogen of the water and set oxygen free.

Effects produced by chlorine's action. — The effect of chlorine varies decidedly with the substance on which it is used. Properly used, chlorine gas is of value for bleaching cotton and linen, but not for whitening silk and wool. Mineral dyes, such as prussian blue, are scarcely affected by it at all. Green trees turn to autumn hues wherever the gas escapes among them in any quantity, for all plant life suffers from exposure to chlorine. Flowers lose their brilliancy under such circumstances, for the gas turns them to a dirty white. On the human frame chlorine exerts an extremely injurious effect. Its action on the water of the mucous membrane is very irritating, while its nauseating odor is hard to bear. It should always be generated under a hood and kept as carefully as possible from the outer air.

Bleaching powder. — The difficulty of generating chlorine directly or of transporting it as a gas makes it essential to have some easy means by which it can be manufactured in some central place, and shipped in portable form wherever needed. Such a means is found in chloride of lime. Chlorine is made in huge generators, and the gas passed over great layers of moist slaked lime. As the gas passes over the lime about half of its weight is absorbed by the solid, forming what is known as chloride of lime. The solid powder may be then shipped in barrels, the lime acting as a holder of the chlorine.

Bleaching powder decomposes when kept for any length of time exposed to the air, and changes to a different chloride, releasing oxygen. When this powder is used for bleaching its chlorine is freed either by the addition of a weak acid or by the immersing in the solution of cloth which has been dipped in acid.

Occurrence of hydrochloric acid. — Hydrochloric acid, the acid from which chlorine is prepared, is one of the great acids of the chemical world. Under its common name, muriatic acid, it is used in thousands of factories, and in hundreds of processes. It is for the most part a laboratory-made acid, but is found in those mighty laboratories of nature, the volcanoes, where in great bubbling pots all sorts of compounds are torn apart and put together. It is found in streams of water which cross volcanic districts and which take the acid in solution from the crater laboratories. For laboratory use it is always prepared from common salt and sulphuric acid.

Making hydrochloric acid. — Common salt is sodium chloride (NaCl). Sulphuric acid is hydrogen sulphate (H_2SO_4). When these substances are brought together and heated, the hydrogen leaves its work of acidifying sulphuric acid and takes up the new work of acidifying the chlorine of the sodium chloride. The acid with which we start, sulphuric acid, disappears, and in its stead comes off a new acid, this time a gaseous one, hydrochloric acid gas. When a mixture of salt and sulphuric acid is heated strongly, this action results:

$$\underset{\text{(salt)}}{\text{Sodium chloride}} + \underset{\text{(sulphuric acid)}}{\text{hydrogen sulphate}} =$$

$$\underset{\text{(hydrochloric acid)}}{\text{hydrogen chloride}} + \text{sodium sulphate}$$

$$2NaCl + H_2SO_4 = 2HCl + Na_2SO_4.$$

When the mixture is heated less strongly all the hydrogen does not go into the hydrochloric acid, but some of it goes into the sodium sulphate, forming a compound known as sodium hydrogen sulphate. This is the white substance generally left in the flask when hydrochloric acid is made in the laboratory. The action under these circumstances might be expressed thus:

Sodium chloride + hydrogen sulphate = sodium hydrogen sulphate + hydrogen chloride.

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$

Properties of hydrochloric acid gas. — The pure gas itself is colorless, the fumes which we see as the vapor rises being liquid acid (or acid in solution), as the steam which shows at the kettle's mouth is liquid water. The yellow color of the solution is due to the dissolved impurities. The gas is heavier than air, can be collected by downward displacement, and has an extremely disagreeable odor, not so bad perhaps as chlorine, but yet evil smelling and dangerous if inhaled in any quantity. Its most striking property is probably its great attraction for water. Hydrochloric acid as it comes off from the generator is a gas. The acid as we use it commonly in laboratory or factory is a liquid. That liquid is simply a solution of the gas. You know how swiftly blotting paper takes up ink, how slowly glazed papers absorb it. Just as those solids vary in absorbing liquids, so liquids vary greatly in their power to take up gases, or, stating it the other way round, gases vary in regard to their solubility in water. Hydrogen, oxygen and air, all these, as we have seen, are practically insoluble in water. Chlorine is slightly soluble. Hydrochloric acid gas is extremely soluble.

Solubility of hydrochloric acid gas. — The fisherman in the Arabian Nights, who opened the sealed jar wherein lay prisoned the genie, saw him rise higher and higher till he towered far above his frightened finder. By shrewd wit the fisherman persuaded the giant once more to reduce his tremendous size and prison himself in the jar. Suppose the genie made himself a hundred times greater than the jar from which he came, he yet would do only one fifth the magical trick that hydrochloric acid gas can do. At freezing temperature and ordinary air pressure, about five hundred volumes of hydrochloric acid gas dissolve in one volume of water. Imagine five hundred quart bottles of gas standing in a row on one side, and one quart bottle of water on the other. The five hundred bottles of gas will dissolve in the one bottle of water. This is a case of solution which goes so far beyond our ordinary ideas of solubility that it seems almost a fairy tale. But it is true.

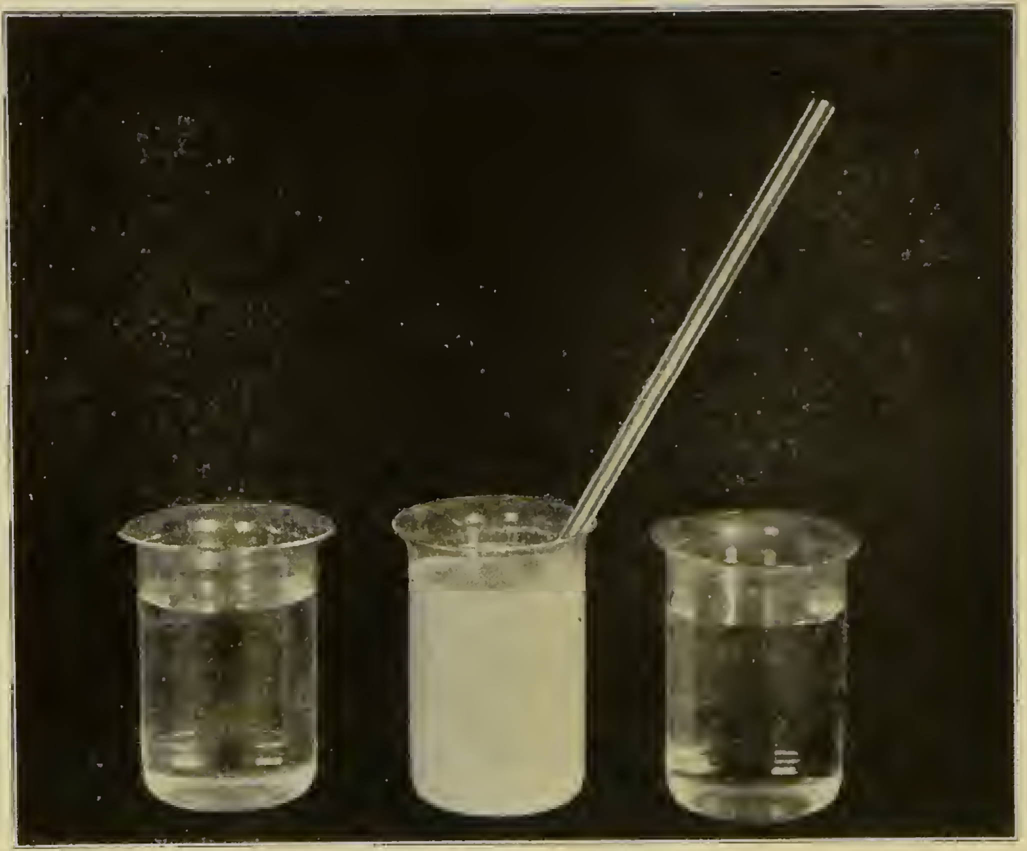
Preparing hydrochloric acid on a large scale. — The making of the thousands of tons of hydrochloric acid that is used in commercial work is accomplished in large part by the same method that was given above for its preparation in the laboratory. So if we say that the laboratory process on a large scale gives the manufacturing process, we shall have the essential facts. Very large quantities of the acid are also produced as a by-product by a process known as the Le Blanc method of making sodium carbonate, a process which we shall discuss fully in our study of the carbonates.

Testing hydrochloric acid.—The common test for hydrochloric acid is a striking one. If we add a solution of a compound of silver, silver nitrate, to a solution of the acid, a soft cloud of curdy white solids known as a precipitate falls which turns purple as the light strikes it. That precipitate is silver chloride. The same cloud forms when we run the acid into a solution of silver nitrate. We may express this by an equation thus:

Hydrogen chloride + silver nitrate =
(hydrochloric acid)
silver chloride
(the white precipitate) + hydrogen nitrate

$$HCl + AgNO_3 = AgCl + HNO_3.$$

Precipitation and filtration. — This action illustrates a principle of great importance in chemistry. It commonly happens that when two substances in solution are brought together the resulting chemical change produces an insoluble substance which falls to the bottom of the vessel. Precipitation is difficult to define simply, but fairly easy to recognize. We can be reasonably sure that precipitation occurs when we see a solid come out of a liquid which is not undergoing evaporation. Such action is called precipitation, and the solid so formed is called a precipitate (abbreviated to ppt.). Precipitates may be separated from the liquid in which they rest by filtration. When so separated, the liquid which passes through is called the filtrate.



PRECIPITATION

Precipitation of silver chloride. The beaker on the right contains a clear solution of silver nitrate. That on the left holds a clear solution of sodium chloride. The center beaker shows the moment of precipitation when these two solutions, added together, produce the solid precipitate silver chloride.

Action of hydrochloric acid on litmus. — Moist hydrochloric acid gas turns dry litmus slowly red, wet litmus strongly red. The solution is strongly acid and gives a strong acid test with litmus and all other indicators.

The chlorides. — The group of chlorides which are made from chlorine and hydrochloric acid have many interesting properties. The next chapter tells of them.

CHAPTER IX

THE CHLORIDES

The solids of the sea. — The ever-moving sea bears on its bosom great fleets of mighty ships carrying the stores of nations in their holds. Its waters perform a thousand different functions in aid of life and industry. Yet outside and beyond the fleets which come and go from shore to shore, outside the water and the friendly rain, drawn from the watery reservoirs of the deep, comes the burden of the ocean, the salts which brine the sea.

THE SEA. 1

A busy harbor. The ships filled with mineral wealth are ready to carry their burdens over the world.

Sea salt. — Taste it where you will, in far Japan, midway the English Channel, or where the tide breaks on the icy wharves of cities lying on far northern shores, the sea gives always the same taste, brackish and salt, always contains the same salts when evaporated, salts which are mostly chlorides and of which sodium chloride, or common salt (NaCl) is by far the greatest part. Our ancestors obtained their salt directly from the sea, by boiling down kettles of sea water until the white salt showed at the kettle's base. Low tide on any summer day will show fringes of glistening salt in hollows of the rocks where pools of sea water have evaporated under the heat of the sun. Such a process has been used for hundreds of years to get the coarse powdered salt of the sea, save that, instead of using as formerly the natural basins of the shore, the sea is either turned into shallow excavations on the shore like docks, or into canals which can be closed by locks. There the pools can lie for days evaporating by the slow heat of the sun, or else be run into kettles, there to be boiled down to a solid residue.

THE SEA. II
Filled with the salts that brine its waters.

The chlorides and sulphates of the ocean. — Sodium chloride, the common salt which is present in sea water, is not the only substance in solution there. In that mixture one finds a whole series of other salts, chlorides of the metals, of potassium, of magnesium, sulphates of magnesium and calcium, even a trace of compounds of gold and other substances. Putting it into English figures, a gallon of sea water holds about 2500 grains of various salts, of which about 1800 are common salt.

Saline springs. — While they can be separated, these various compounds cannot be taken apart easily and economically, so that the salt which we use on the table comes mostly from other sources. Every reader of Indian stories, every lover of Cooper's Leatherstocking Tales, has read of the salt licks where the deer came down to get the salt they craved. Those salt licks were the soil about springs which held salt in solution, and which, bubbling through the earth, bore with them salt to be deposited on the soil. Some American salt comes from such saline springs, and much of it comes from near Cooper's old country in central New York. Abroad, as here, salt comes in great quantity from rock salt deposits, in some of which beautiful crystals are found, which are like clear-cut colorless gems. Absolutely pure salt is not commonly found on our table, as is shown by the moistening of this compound in summer from the absorption of the water vapor of the air. Pure salt absorbs very little water in this way. It is the impurities present, such as magnesium and calcium chlorides, which attract the water and moisten the salt.

The family of the chlorides. — Few studies are more fascinating than the careful inspection of a single group of compounds, all made in a similar way, all bearing marked family resemblances, yet, in their different action on differing kinds of human endeavor, so far apart as to be almost diametrically opposed. There are many sets of compounds of this type beside the chlorides; sulphates from sulphuric acid, nitrates from nitric acid, acetates from acetic acid, sulphides, carbonates, oxides; the list can be carried on to no slight length. We cannot study all as groups, but we should recognize that these groups form families and communities which are, one might almost say, landmarks of the chemical world. When chlorine unites directly with a metal, or when hydrochloric acid acts on a metal or a metallic compound, a chloride is commonly produced. Table salt is the commonest chloride we have, the most wide-spread, the most necessary to man, but a brief study of a few of the many other members of this group will repay the time taken to consider them.

Calcium chloride. — From invisible inks to medicines that heal the sick, from soldering tin cans to aiding in fortifications, range the uses of some of the chlorides. The possibilities inherent in this group might be matched and excelled by those possessed by some of the other groups. Calcium chloride ($CaCl_2$) is a deliquescent salt, for example. It eagerly reaches out for water in the atmosphere, seizes it, and from a dry powder turns first to a moist substance of the consistency of sticky mud on a warm March day, and then to a thick solution. When we desire to dry a gas which contains water vapor we pass the gas through a calcium chloride tube, as it is called. It enters charged with water. It emerges dried. The calcium chloride took the aqueous vapor as the gas passed over it. It is important to keep many things from rust, from the action of the damp and humid air. In many cases rust, a type of oxidation, proceeds swiftly in such air, while no oxidation occurs in dry air. In consequence we place calcium chloride in a fine balance, or polished delicate piece of apparatus, to gather all the moisture that penetrates inside the case. In fortifications and on shipboard where gunners were likely to be exposed to fire from bursting shot or shell, or from the blaze of heated wood or metal, great mantlets were formerly used, mats of rope soaked in a solution of calcium chloride. These mantlets never grew dry, never could take fire, for the salt solution with which their fibers were impregnated always kept them moist.

Cobalt chloride. — What reader of Edgar Allen Poe can forget the thrill when the blank sheet whose death's head led to the hidden treasure springs suddenly into written script when held to the flame. Sympathetic or invisible inks, inks invisible in cold, visible when heat is applied, have carried many a conspirator's letter, guarded many a state secret. Cobalt chloride ($CoCl_2$) is one of salts which give such an ink. In dilute solution, when it contains water of crystallization, it is practically colorless. Write a note with a solution of cobalt chloride and the pages seem blank. Hold it to the flame and the blue letters appear as if by magic. The heat drives off the water of crystallization into the air, and leaves the cobalt chloride blue, its normal color when dry. A saturated solution of cobalt chloride is pink, the real color of the salt when water of crystallization is present. Some of the little doll barometers, whose fluffy skirts turn pink when the air is charged with water vapor, as when a storm is near, and blue when fair dry weather comes, owe their powers of prophecy to cobalt chloride. The wet atmosphere fills the cobalt chloride in the cloth of the dress with water of crystallization. It turns pink. A dry atmosphere evaporates the water of crystallization from the toy barometer and makes the blue that indicates a cloudless sky.

Two chlorides of mercury. — Whatever else the old-fashioned country doctor carried in his medicine case, he always took calomel (mercurous chloride, Hg_2Cl_2), containing two atoms of mercury and two of chlorine, along. Whatever other antiseptics the modern surgeon possesses, he is sure to have on his shelf a bottle of corrosive sublimate ($HgCl_2$), which contains one atom of mercury and two of chlorine. In few cases in chemistry do we find two compounds so closely related and yet so entirely opposite in their action on the human frame, so Jekyll-and-Hyde-like, in their character. Calomel, though used to-day far more sparingly than of old, is still a valuable and useful drug, poisonous only in comparatively large quantities. Corrosive sublimate ($HgCl_2$), invaluable as an antiseptic, is one of the most poisonous substances when taken internally that we have. It is a great disinfectant, a most valuable destroyer of insect life, especially of household pests, yet its use must always require great caution. Too much care cannot be taken to prevent its being accidentally swallowed. In case any corrosive sublimate is taken internally, instant emetics and the stomach pump must be used to remove it. White of egg is a valuable antidote, but should not be relied on altogether. Corrosive sublimate is used as a preservative for anatomical preparations, for wood and other substances likely to decay. See what a difference exists in the action of these two substances, both containing mercury and chlorine, the sole difference in their composition being that the corrosive sublimate contains twice the proportion of chlorine that calomel does.

Potassium and magnesium chlorides. — Two of the light metals which seldom occur in their elemental forms, potassium and magnesium, are derived from their chlorides, potassium chloride and magnesium chloride respectively. Potassium chloride (KCl) comes largely from the refuse of sugar factories, where the sweet product is made from the sugar beet. Magnesium chloride ($MgCl_2$) is found commonly in company with sodium chloride in the brackish waters of the sea, in saline springs and in salt mines.

Gold and silver chlorides. — The chlorides of the precious metals, gold and silver, (silver chloride = $AgCl$; gold chloride = $AuCl_3$), both exhibit one property peculiar to themselves. They break down under the compelling power of the sun's rays, giving free gold or silver, and this action, as we shall see later, lies at the basis of the photographic art.

Tin chloride. — The soft rose pink of many a gown or ribbon bow owes its beauty to the chlorides of tin ($SnCl_2$ and $SnCl_4$), of that same tin that we see in a tin plate. Unite that metal with chlorine, form a chloride, and you have a salt which impregnates the fibers of cloth to give some of the most beautiful pinks that the dyer's art can produce.

Zinc chloride. — Passing from ribbons to tin cans we find the next use of the chlorides. Zinc chloride ($ZnCl_2$) is used to cleanse the metal in the soldering of tin cans. This chloride is poisonous, so that careless employees have sometimes caused grave illness by spilling the salt into cans containing preserved fruits and vegetables. Modern methods of sealing cans have largely obviated this trouble.

Ammonium chloride (Sal ammoniac). — The list of the uses of the chlorides might be widely extended: we may cite ammonium chloride (sal ammoniac, NH_4Cl), the white salt which in solution in the so-called "Lelanché battery" rings the electric door-bell; barium chloride ($BaCl_2$), the valuable salt which precipitates the sulphates; copper chloride ($CuCl_2$), which gives the paint known as Brunswick Green; iron perchloride ($FeCl_3$), the medical tincture of iron given to supply iron in the blood; these and many another chloride aid in the arts or in medicines.

Crystals of Ammonium Chloride
Sal-ammoniac.

The preparation and properties of the chlorides. — The chief method of making chlorides is by means of the action of hydrochloric acid. A large number of the metals dissolve in this acid, leaving as a result a chloride in solution. In consequence we have chlorides of almost every metal, and the constitution of almost every one is simply the sum of the metals and chlorine. Such substances are called metallic chlorides, so, as well as sodium chloride, we have gold chloride, silver chloride, mercury chloride, lead chloride, and so on. Most of these are white powders, some are crystals, and all but three, the chlorides of lead, silver, and mercurous mercury, are soluble in water or dilute acids.

Reactions for forming chlorides. — The general method of forming chlorides by the action of hydrochloric acid may be illustrated by the way zinc chloride is formed when hydrogen is made.

Zinc + hydrogen chloride = zinc chloride + hydrogen.

$$Zn + 2HCl = ZnCl_2 + H_2$$

When an oxide of a metal is attacked by hydrochloric acid, the action is similar, but water is formed instead of hydrogen.

Zinc oxide + hydrogen chloride = zinc chloride + hydrogen oxide (water).

$$ZnO + 2HCl = ZnCl_2 + H_2O.$$

When the salt of another acid like a sulphate is used, another acid made up of the compound is generally formed. For example:

Zinc chloride + hydrogen sulphate (sulphuric acid) = zinc sulphate + hydrogen chloride (hydrochloric acid).

$$ZnCl_2 + H_2SO_4 = ZnSO_4 + 2HCl.$$

When we add hydrochloric acid to the hydroxide of a metal, we have a case of acid + base = salt + water.

Zinc hydroxide + hydrogen chloride = zinc chloride + hydrogen hydroxide (water).

$$Zn(OH)_2 + 2HCl = ZnCl_2 + 2H_2O\ (2H{\cdot}OH).$$

(Though we can write the equation in this way, it is worth while to remember that water very seldom acts as a hydroxide.)

The general way to proceed if we want to make a chloride from any substance is to act upon the substance with hydrochloric acid.

The direct formation of chlorides from chlorine. — Chlorides are sometimes made by direct action of chlorine on various elements. Sodium, antimony, and copper all unite directly with chlorine, and several other elements act in a similar fashion. A representative equation of this group would be:

Sodium + chlorine = sodium chloride.

$$Na + Cl = NaCl.$$

Copper + chlorine = copper chloride.

$$Cu + Cl_2 = CuCl_2.$$

Testing for chlorides. — The chief test for the chlorides is the same as that for hydrochloric acid, the addition of silver nitrate. This test is almost universal, and may be exemplified by the action of silver nitrate on sodium chloride.

Silver nitrate + sodium chloride = silver chloride + sodium nitrate.

$$AgNO_3 + NaCl = NaNO_3 + AgCl.$$

The silver nitrate which forms is a white cloudy precipitate, which turns purple in the light.

The silver group of insoluble chlorides. — This action of silver nitrate illustrates the general rule which applies to all three of the so-called group of insoluble chlorides or silver group, which include three metals, silver, lead, and mercurous mercury. In almost every compound of these three, the addition of a chloride, such as hydrochloric acid or sodium chloride, produces a white precipitate. With silver, silver chloride is formed; with lead, lead chloride; with mercurous mercury, mercurous chloride.

Formation of lead chloride:

$$Pb(NO_3)_2 + 2HCl = PbCl_2 + 2HNO_3.$$

Formation of mercurous chloride:

$$Hg_2(NO_3)_2 + 2HCl = Hg_2Cl_2 + 2HNO_3.$$

CHAPTER X

FLUORINE, BROMINE, AND IODINE

Symbol for fluorine (F), for bromine (Br), for iodine (I)

The halogen family. — The most fascinating fairy stories that we possess, commonly begin with four brothers, starting out into the world to seek their fortunes. One of the brothers may be light, another dark, one haughty, another humble, but all possess certain family characteristics. When these brothers come from the realm of the sea, their story becomes even more fascinating than otherwise, and that is the realm from which appear the family of the halogens, fluorine, chlorine, bromine, and iodine, whose very name, halogen, comes from the Greek ἅλς (the sea), and whose sodium salts one and all resemble sea salt. These four substances are the most closely related of any of the elements, and in their relation show as interesting differences and likenesses as a family of human brothers might reveal.

Chemical families in general. — The halogens are by no means the only family group in chemistry. Far from it. Every element is related more or less closely to certain other elements, which together form a group or family with certain characteristics in common. But none of these families show their resemblances so strongly as the halogens. None are so strongly marked as the halides, as the individual members of this family have sometimes been called.

Fluorine, chlorine, bromine, iodine. — Chlorine, the second and most important of the group, has already been considered. Fluorine, bromine, and iodine, the other members, will first be considered briefly by themselves, and afterward we may consider the family as a whole.

The action of fluorine. — Of all the halogens, fluorine is the most active. The lines which cut across a graduated glass, the letters etched into a bottle's side, were made by fluorine, or by its nearest compound, hydrofluoric acid. This element eats through glass by dissolving the silica which it contains, and silica is sand. Think of the insoluble sandy bed of the ocean and you can realize how powerful this agent is. The gas fluorine leaps eagerly into combination with almost every metal forming fluorides. Gold and platinum alone of the ordinary metals escape. Hydrogen, bromine, iodine, sulphur, phosphorus, carbon and other elements burst into flame in this gas. Water breaks down in its presence, owing to the great attraction between this element and the hydrogen of the water, while fluorine readily decomposes substances like coal gas called hydrocarbons.

Fluorspar and hydrofluoric acid. — Few compounds show a more beautiful form than fluorspar, the fluoride of calcium (CaF_2), which is the commonest compound in which fluorine is found. In large masses of cut crystals it shines with wonderful shadings, its colors passing from deep violet through royal purple to azure blue and rich soft greens. Whether colored or colorless few minerals exact more attention than this compound of calcium, the metal of lime, and fluorine. When strong sulphuric acid is added to calcium fluoride and the mixture is heated in a leaden dish, hydrofluoric acid ($CaF_2 + H_2SO_4 = CaSO_4 + 2HF$) or hydrogen fluoride is evolved. This acid resembles hydrochloric acid in many ways, but is far more powerful. Its smell is more pungent and irritating. Dropped on the hand, it causes painful sores. This acid combines with water with such eagerness that it hisses like red-hot iron when it strikes the cold liquid. The solution obtained from such combination, however, is not a compound of hydrofluoric acid and the water. It is simply a solution of hydrofluoric acid in water as our ordinary hydrochloric acid is a solution of hydrochloric acid in water. When hydrofluoric acid gas is used for etching, the glass is coated with wax, and the lines to be attacked are scratched through the wax uncovering the glass. The glass is then placed where the acid gas can attack it. The acid eats away the glass where it is exposed by the scratches, but cannot attack the surface covered with the wax. By this action a permanent rough outline is left which stands out against the smooth polish of the glass. This forms a durable etching. Glass is made up largely of sand (silicon dioxide SiO_2), and its silica enters into combination with the fluorine of the hydrofluoric acid.

The production of fluorine from hydrofluoric acid. — We think of the action of fluorine largely in terms of its acid, because of the fact that it is so difficult to obtain the element free. Only within the last twenty-five years has this substance been obtained pure and, when the final separation was accomplished, it was done by breaking up a solution of acid potassium fluoride, HF, KF, in anhydrous hydrofluoric acid, HF, with the electric current. The separation was finally accomplished by Moissan, the great French chemist, and was done in tubes of platinum mixed with the rare metal iridium. When the current is passed through this solution, hydrogen is evolved at one pole and fluorine at the other. The process is not a simple separation of the hydrofluoric acid, but also involves the potassium present.

The capture of fluorine. — The whole scientific world watched with interest this experiment to gain an element which had been known to exist for decades, but which could never be prisoned long enough to be examined. Like a criminal, so wise and crafty that he could never be held by lock or bolt, fluorine, until Moissan captured and examined it, could never be obtained. It could not be held in ordinary tubes of glass or metal, for it dissolved them swiftly, escaping or turning into different forms. Stoppers of wood or cork would not confine it. It seemed as though it would escape from every bond. Only when the rarest and most resistant metals, like platinum and iridium, were used for containers and transmitters, when stoppers were made of fluorides, which had already taken up all the fluorine they could hold, and transmitting tubes were constructed of rare and valuable platinum, was the element fluorine separated and examined.

Properties of fluorine. — The gas fluorine which rises from the action of the electric current has a sharp pungent odor, a property characteristic of every one of the halides. Its color is greenish yellow, but of a lighter shade than the yellow green of chlorine, and even more yellowish. It is heavier than air and falls when collected to the bottom of a receiver, while at a very low temperature it becomes a pale yellow liquid. This gas is the first and strongest member of the family. Its weight is the least. Fluorine and hydrofluoric acid may both be tested by their power of dissolving glass.

Bromine. — Though chlorine comes naturally second in the list of four, we will give place here to bromine, since we considered chlorine in the chapter preceding this. Bromine, like chlorine, comes from the sea. It is almost always found mixed with some compounds of its nearest relative, and, as salt is found in saline springs, so are bromides or compounds of bromine. This element varies materially from both fluorine and chlorine. The two former are energetic gases. This element is the only elementary liquid known at ordinary temperatures save mercury. It is the only non-metallic element which appears in a liquid form. More than all else, it is characterized by its intensely trying and irritating odor. Its very name comes from the Greek *βρῶμος*, a stench, and its odor does not belie its name. In appearance it is a dark, reddish brown liquid, which at ordinary temperatures is about three times as heavy as water. It is soluble in water, and easily soluble in various other reagents. It boils far below the boiling-point of water, and its vapor, like the liquid, is extremely poisonous and burns the flesh badly. In the laboratory bromine is made by heating one of its salts, the bromide of potassium, with sulphuric acid and with manganese dioxide (the same substance which we use in making oxygen and hydrochloric acid). ($2KBr + MnO_2 + 2H_2SO_4 = MnSO_4 + K_2SO_4 + 2H_2O + Br_2$.) When the sulphate combines with the potassium the oxygen of the manganese forms water and bromine comes off as a heavy brown gas. The action is quite similar to that which produces chlorine. Its preparation should be attempted only with care and under a hood. Most of the bromine that we use in commerce comes from the liquor left from the saline springs of the salt licks which give common salt. Like fluorine and chlorine, bromine has its acid, hydrobromic acid (HBr). This colorless gas, with a strong irritating smell, dissolves freely in water, giving a solution entirely similar to hydrochloric and hydrofluoric acids. It may be tested by the dark, red, brown color which it imparts to the organic carbon bisulphide.

Some uses of bromine. — The pink string that ties the drug store bundle or the Christmas present, the red ink that lines the ledger or corrects the theme, formerly came from bromine, used to make a dye which tints these substances. (To-day red aniline dyes have largely taken the place of bromine for these purposes.) Bromides are used largely in medicine, while every photographic house advertises bromide papers. In this respect bromine acts only as an aid to gold and silver, of whose work in photography we shall speak later.

Iodine. — The last of the family is the heaviest in weight and is the only solid of the four, iodine. A striking advance in weight appears as the family proceeds: fluorine first, a lighter gas than chlorine, then chlorine, a heavy gas, then bromine, a liquid, and now iodine, a solid. Free iodine is never found in nature, any more than the other halogens, since all these halides combine too readily to be found free. But the element exists in many places in the natural world. The watercourse of the brook, the fertilizer of the farm, the sponge of the deep sea, sometimes contain this element, but most of all it comes from that deep sea tangle, the kelp of the shore.

The discovery of iodine. — The story of the discovery of iodine reads like a fragment from a fairy tale. Almost a century ago, a French investigator, Courtois, was heating liquids obtained from the kelp which washes to and fro on the northern coasts of France. He was striving to find niter, to separate out the nitrate crystals which have fertilizing value. As he watched, the liquids boiled quietly on, grew lower, and, as he looked away, grew lower still. Suddenly he turned to see a spreading vapor of rich and royal purple rising from his rude vessels and filling the room with violet clouds. That vapor was iodine. Courtois seized the clue, ran down its course, and brought the element to the knowledge of the world. To-day the seaweed which first opened the possibilities of a new element to its discoverer's eyes still furnishes a part of the world's supply of iodine, though more, probably, comes from iodides found in one of the compounds Courtois was seeking to obtain, sodium nitrate. This substance as collected contains a small amount of sodium iodide (the iodine salt of sodium) which comes out when sodium nitrate is finally cleansed. Preparation of iodine is like that of bromine, $2KI + MnO_2 + 2H_2SO_4 = K_2SO_4 + MnSO_4 + 2H_2O + I_2$.

The properties of iodine. — Iodine in its natural condition is a blue-black lustrous solid which crystallizes in shining leafy plates. It melts at a point just above the boiling-point of water, and turns into the beautiful violet vapor which led to its discovery. Indeed, it vaporizes at ordinary temperatures, and the bottle wherein iodine crystals are kept commonly seems darkened by the vapor within. When iodine vapor is heated strongly it turns to a deep indigo blue.

Sublimation. — The change which takes place when solid iodine changes to gaseous iodine differs materially from that which goes on when most solids change to gases. The normal way in which most substances change their state of matter is to pass from solid to liquid, then from liquid to gas, as ice changes to water, water to steam. In the case of a few substances, notably iodine and ammonium chloride (the sal ammoniac of batteries which has already been mentioned), the intermediate stage of liquid matter cannot be seen; the gas appears to turn directly into a solid and vice versa.

Such a change is called sublimation. Iodine which changes from a gas into a solid without apparently passing through a liquid state is said to sublime.

Solubility of iodine. The starch test for iodine. — Iodine is slightly soluble in water, and freely soluble in alcohol. Like the other halogens it forms an acid with hydrogen, hydriodic acid (hydrogen iodide (HI)), which, acting on metals, forms compounds called iodides. None of the reactions of this element are more interesting than its power to form a strong blue color when brought in contact with starch. This test, the formation of a blue coloration when starch and iodine meet, is at once a test for iodine and a test for starch. Since starch is one of the most widely distributed substances of the organic world, the value of this test is great. Iodine is chiefly used in medicines as a counter irritant.

The similarities of the halogens. — The four members of the halogen family are fluorine, chlorine, bromine, and iodine. We have seen how they differ, what their personal peculiarities are: let us see now how they are alike. Every one of them unites with hydrogen in the proportion of one to one, to form a series of compounds remarkably similar in their properties. Fluorides, chlorides, bromides, iodides, each of these groups have a combining power equal to that of hydrogen. Every one of the elements is a substance of active combining powers. When we come to examine their exact condition in this respect, we find some surprising facts. The energy with which they combine is in an inverse ratio to their weights, taken with respect to hydrogen, the standard. Roughly speaking, that statement means just this: The lighter the halogen, the more eagerly it combines with other substances. The heavier it is, the slower it is in combination. Fluorine, the lightest of the group, is intensely energetic, eats glass away, unites explosively with hydrogen in the dark to form hydrofluoric acid (HF), and cannot be held in ordinary receivers because it combines with the material from which they are made. Chlorine, a heavier gas, unites explosively with hydrogen in the light, forming hydrochloric acid (HCl), forms compounds readily enough, yet does not exhibit the fierce energy of fluorine. Bromine, a fairly heavy liquid, forms hydrobromic acid (HBr) only when in contact with hydrogen in a red-hot tube, uniting in this way with less ease than chlorine forms hydrogen chloride. Iodine, the solid, forms hydriodic acid (HI) directly only by aid of the catalytic action of finely divided platinum, while the iodides show but little of the fierceness of combination which the other halogens show. Iodine in iodides can be replaced by bromine. Bromine in bromides can be replaced by chlorine, and chlorine in chlorides can be replaced by fluorine. It almost seems as if the heavier the element became, the less its power of accomplishment.

Mendeléeff's periodic law. — After all, this story of the four brothers shows one thing, the relationship that can exist between elements. If we went into the matter more deeply we should find a law, Mendeléeff's periodic law, which shows stranger relationships even than are shown here, and by which tables may be prepared, placing every element in its proper place among its own family. Our own study of the matter, however, must stop with the recognition that, like the separation of the human race into families and races, so in chemistry we may find groups and divisions among the elements in which every member bears a certain relationship to the others.

CHAPTER XI

CARBON DIOXIDE

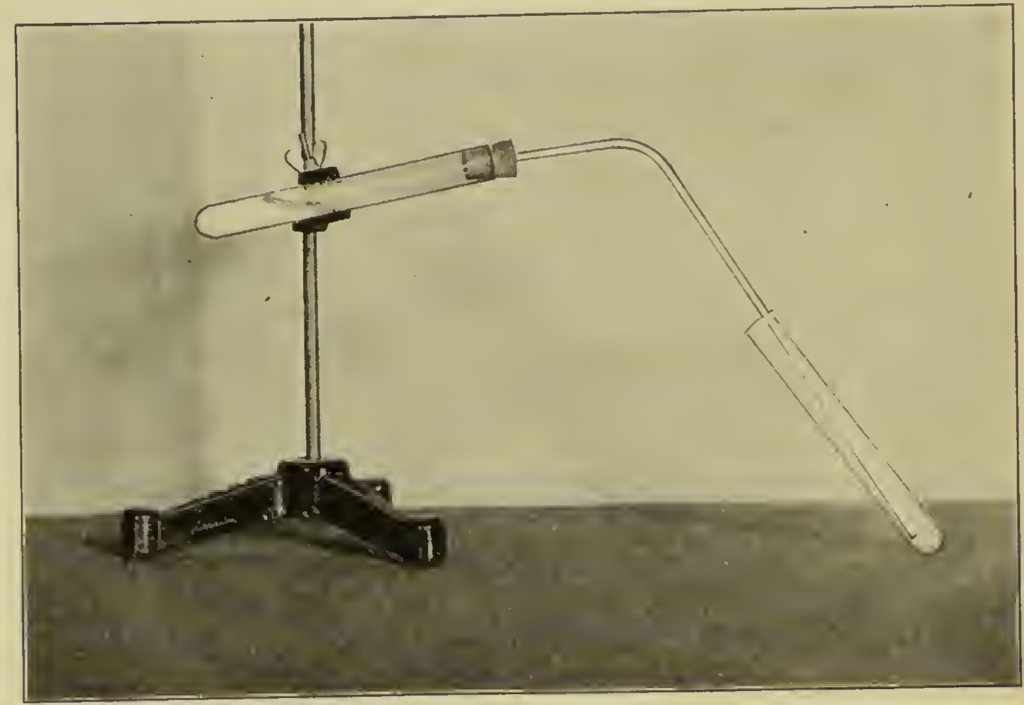
Formula for carbon dioxide = CO_2

Occurrence of carbon dioxide. — Carbon dioxide is the chief product formed by fire or decay, the main food of the plants, and, if we bar water vapor, the most widely distributed gaseous compound of the whole earth. Carbon dioxide exists in the air of mid-ocean and of the prairie. It is found in comparatively large quantities in the atmosphere of cities, where factory chimneys pour out hot gases from coal fires. It comes from every hearth fire, from every stove, from every furnace. It is produced by the dying flower or leaf, by the decaying tree and shrub. It occurs in every breath sent forth by a human being.

Carbon in the world of nature. — The reason for this great production of carbon dioxide is not hard to find. Carbon is present in all organic substances. Wood and coal and vegetable matters generally belong in this organic class. When these substances burn or decay the carbon in them oxidizes and uniting with oxygen forms carbon dioxide, commonly called carbonic acid gas. The human body is largely made up of carbon. Our food is chiefly carbon, our stock of it being constantly replenished with every meal, since our store of the element is continually burned to furnish us with the necessary heat for life, and for the energy which we need for our daily tasks. One product of our body burning is carbon dioxide, which comes out from the lungs with the breath and enters the air.

Production of carbon dioxide by roasting calcium carbonate. — Probably most of the carbon dioxide which is produced commercially comes from the roasting of carbonates. These bodies, as we shall see later, are the salts of carbonic acid, and they readily part with carbon dioxide when heated to a moderately high temperature. Calcium carbonate ($CaCO_3$) is the substance most commonly employed for this purpose, for this carbonate, when heated strongly, breaks down, leaving solid calcium oxide behind, and setting free carbon dioxide gas. The reaction for roasted calcium carbonate is $CaCO_3$ (heated) $= CaO + CO_2$.

Production of carbon dioxide by the action of an acid on a carbonate. — While calcium carbonate can be roasted in the laboratory, yet a simpler method gives better satisfaction. This method, already considered in our study of acid, base, and salt, consists of adding an acid to a carbonate. It is employed to give carbon dioxide gas whenever the two powders of a



Making Carbon Dioxide

Carbon dioxide is generated when powdered marble and hydrochloric acid are simply mixed together. The action proceeds without the aid of heat. The gas is collected by downward displacement.

seidlitz powder are put into solution and mixed together. The gas which foams up in the glass is carbon dioxide, which has been produced by the addition of an acid substance to a carbonate. With but very few exceptions this principle holds good. When an acid attacks a carbonate, carbon dioxide is evolved.

In the usual laboratory method of carrying out this process hydrochloric acid (hydrogen chloride) is added to calcium carbonate, and calcium chloride, hydrogen oxide (water) and carbon dioxide gas are the products ($CaCO_3 + 2HCl = CaCl_2 + CO_2 + H_2O$). The following paragraph explains this reaction.

Carbonic acid. — One would naturally expect that carbonic acid, hydrogen carbonate, H_2CO_3, would be one of the final products of this action, instead of carbon dioxide gas and water, and that the reaction just given would read $CaCO_3 + 2HCl = CaCl_2 + H_2CO_3$, but this acid is so weak that it invariably breaks down with evolution of carbon dioxide gas whenever it is formed ($H_2CO_3 = H_2O + CO_2$). This important fact plays a part in many a reaction.

The reaction which takes place between sulphuric acid and a carbonate. — When sulphuric acid (hydrogen sulphate) is added to calcium carbonate in the form of lumps of marble carbon dioxide comes off at first, but ceases so soon that very little of the gas can be obtained by this process. This is due to the insolubility of the calcium sulphate formed, which coats the calcium carbonate and shields it from further action of the acid. When finely powdered chalk (also calcium carbonate) is used this difficulty does not occur to a great extent.

If we write these actions in equation form they would read as follows: (Calcium sulphate is plaster of Paris).

Calcium carbonate + hydrogen sulphate = calcium sulphate + hydrogen oxide + carbon dioxide.

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2.$$

The weight of carbon dioxide gas. — Few of the gases compare with carbon dioxide in regard to weight. It is more than half again as heavy as air, and can be poured like water from one beaker to another. It is easily soluble, invisible, odorless, and tasteless. According to some physiologists it is dangerous to health and life when found in any large proportion in the air, stopping the escape of the carbon dioxide from the lungs and preventing the taking in of the oxygen of the air necessary for the body changes. As every man, woman, and child is a factory of carbon dioxide running night and day, the danger of close or crowded rooms should be self-evident. The necessity for opening windows is equally evident, whether the carbon dioxide be as poisonous as some think or not, since the closeness of any room, caused by exhaustion of the oxygen by humidity or the crowding together of many persons, shows in the rise of the per cent of carbon dioxide in the air. Close rooms are one of the dangers against which we must constantly struggle.

Carbon dioxide gas in nature. The "cycle of carbon." — When we think of these things or look on the many chimneys of the town pouring out carbon dioxide in the air the question naturally must arise, "With all this production of carbonic acid gas, how can the atmosphere keep from acquiring so much of the gas as to be dangerous to mankind?" No small part of this gas is swept through the heavens by the blowing winds, but one answer to this question is found in the fact that carbon dioxide is the food of all plant life. Every green leaf, every flowering shrub, is constantly absorbing carbon dioxide, constantly clearing the air by that absorption. This gas is found at the end as it is at the beginning of plant life, for it is produced by vegetable decay. It starts as the food of the green world about us, and it ends as the product of its decomposition: so passing from birth to death, it passes through a cycle, a round of changes known as the "Cycle of Carbon." This cycle begins with the carbon dioxide of the air. Every leafy cell is a chemical factory endowed with such intricate machinery that it can use the carbon dioxide of the air in its growth, and break it down into its component parts. So doing it is able to retain the carbon as a part of its structure and set free the oxygen to go forth once more into the atmosphere. All plant life is constantly separating this compound gas in this manner, and building the carbon of the gas into its own structure, thereby to grow and increase. But the life of any plant is but a little while. If it serves for food for man or beast its carbon enters another body frame, is there burned for fuel and comes forth once again in the form of carbon dioxide, to be once more taken in by the plant life to serve as food. Even if the plant gives no support for man or beast, it must in time decay, and from this decay comes forth the carbon again in the form of carbon dioxide. Other substances, compounds of carbon, may come as well, but the chief form found when plant life decays or burns is this same oxide of carbon. So turns the great cycle of carbon, first in gaseous oxide form entering the plant, remaining in plant life or animal for its allotted time, and finally passing back to the original gaseous form.

Carbon dioxide in soda water. — The cycle of carbon shows the use of carbon dioxide in nature as food for the plants. Its solubility affords a field for its use commercially. An important use for the gas is found in its presence in natural mineral waters, but more important yet is its use in soda, the ordinary variety of the soda fountain, in ginger ale, and all similar drinks. Soda is made by simply charging water with carbon dioxide gas under pressure, and then sealing it to keep it from escaping into the air; that is, soda is simply a solution of carbon dioxide gas. Place this solution in a siphon or in a closed tank and the water holds the gas fast, but once open the gate to the atmosphere and the excess of carbon dioxide instantly rushes forth with the bubbling effervescence of a long glass of soda water.

Carbon dioxide in fire extinguishers. — Fire extinguishers furnish another commercial use for the gas. Since carbon dioxide has already taken up all the oxygen it can hold, oxidation in its presence is an impossibility. We know that fire of the ordinary kind comes only when substances have free access to oxygen, so that we may readily understand that the presence of carbon dioxide tends to put out a fire. Burning would be as impossible in an atmosphere of carbon dioxide as in one of nitrogen. Not only does it diminish the amount of oxygen in the air, but, being a very heavy gas, so heavy indeed that it can be poured almost as a liquid, its very weight acts like a blanket covering the flames. Taking advantage of this peculiarity, fire extinguishers are made which contain chemicals which added together form carbon dioxide. Sodium carbonate and sulphuric acid are generally used in such extinguishers, the sulphuric acid being placed in a bottle inside a layer of the carbonate. When the extinguisher is tipped upside down the acid runs out on to the carbonate. The carbon dioxide which is evolved escapes through a tube.

Testing for carbon dioxide. — No gas has a more characteristic test than has carbon dioxide. If we take a solution of lime water (the common lime water of the drug store), which, chemically speaking, is calcium hydroxide ($Ca(OH)_2$), and pass this gas into it, either from a generator or by passing our breath into the solution through a tube, we soon see a white precipitate forming throughout the liquid. That precipitate is calcium carbonate. It has been formed in this way. Calcium hydroxide is made up by adding water to calcium oxide. When carbon dioxide bubbles through it the water drops off and the gas is added directly to the oxide, forming calcium carbonate ($Ca(OH)_2 + CO_2 = CaCO_3 + H_2O$).

This reaction may be considered as a subtraction followed by an addition. The subtraction would consist of the loss of a molecule of water from the calcium hydroxide $Ca(OH)_2 - H_2O = CaO$. The addition would consist of adding a molecule of carbon dioxide to the calcium oxide $CaO + CO_2 = CaCO_3$. Together this addition and subtraction would form the reaction just shown.

Carbon dioxide in bread-making. — The fascination of the rising dough, the marvel of the swelling pans of biscuit, never ceases to interest the wondering child. Why does the dough rise in the bread raiser? Whence came the pores in the firm white mixture of well-kneaded bread? The swelling of the dough came from the gases imprisoned within, and gaseous carbon dioxide, which causes the rising of the bread, is produced by the fermentation of sugar with yeast. The raising of soda biscuit is done by the same means, save that instead of taking yeast and allowing it to ferment the organic substances present, baking soda (the household sodium bicarbonate) is taken and added to some substances by which carbon dioxide may be evolved. Thereupon the soda biscuits rise from the gas within. In all fermentation processes carbon dioxide is evolved.

Carbon monoxide. — The blue flame flickering on the surface of an open coal fire is carbon monoxide (CO), the second of the oxides of carbon, called monoxide because it contains but one atom of oxygen. In the lower part of a fire carbon dioxide is formed when the air comes in contact with the coal at the bottom of the grate. The gas, bearing two parts of the oxygen as it rises, meets the red-hot carbon above and yields to it one part of its oxygen. Carbon monoxide is formed which, burning with a blue flame, turns back once more to carbon dioxide ($2CO + O_2 = 2CO_2$).

The reduction of iron ores. Water gas. — The burning of this gas, the change from monoxide to dioxide, gives out tremendous heat, and from the stores of energy thus sent forth we get two uses of great value. First, there is its use in the smelting of iron and other ores. In this the carbon monoxide, seeking a second oxygen atom, takes from the ores their combined oxygen and leaves them free. Second, there is the use of carbon monoxide in water gas. It will be remembered that when steam was passed over red-hot iron, iron oxide was left and hydrogen evolved. In similar fashion, when steam is passed over red-hot carbon, carbon monoxide and hydrogen are the products, a mixture known as water gas ($H_2O + C =$

$CO + H_2$). Both of these gases burn with feeble light, but with large output of heat. Heat is the chief requisite in gas to-day, and water gas produces heat from the oxidation of carbon monoxide to the dioxide, and of hydrogen to hydrogen dioxide, water. Neither of these gases give off much light, but, by adding a small percentage of a naphtha gas, whose gas gives visible flame, to them, we get luminosity as well as heat.

Danger in the general use of water gas. — Water gas, however, is not generally considered to be as available for general use as coal gas, on account of its heavy percentage of carbon monoxide, for this gas, when unburned, is generally considered to be a direct poison, acting upon the blood in a rapid and deadly fashion.

Preparation of carbon monoxide gas. — Carbon monoxide is prepared in the laboratory by acting upon oxalic acid with strong sulphuric acid. When this is done carbon dioxide, carbon monoxide, and water are evolved. The sulphuric acid, by reason of its affinity for water, takes up the water, the gases pass off, and the carbon dioxide is absorbed by a solution of sodium or calcium carbonate, leaving the carbon monoxide free. This gas is tested by the fact that it burns with a blue flame, forming carbon dioxide. In preparing carbon monoxide from oxalic acid by the use of sulphuric acid the reaction may be regarded as a subtraction of a molecule of water.

Formula for oxalic acid = $C_2H_2O_4$.

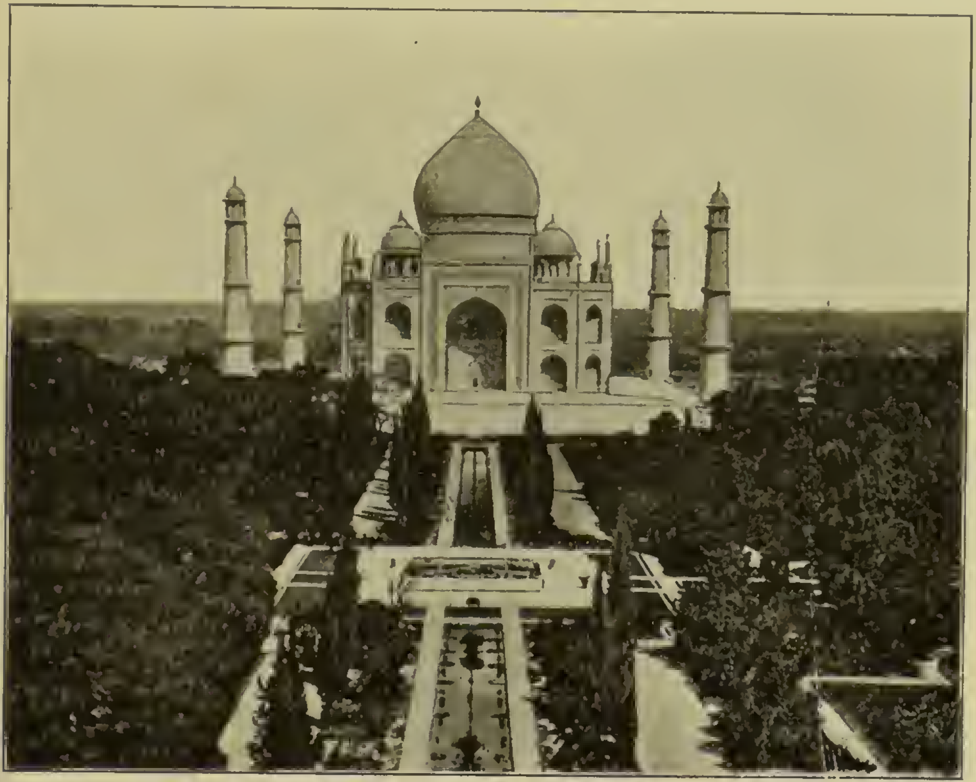
$$C_2H_2O_4 - H_2O = CO_2 + CO.$$

Apparently we might suppose that this subtraction of a molecule of water would leave a compound with the formula $C_2O_3(C_2H_2O_4 - H_2O = C_2O_3)$. Practically no such compound exists. We obtain two compounds instead of one, CO_2 and CO, which together contain two atoms of carbon and three of oxygen.

CHAPTER XII

THE CARBONATES

Occurrence of the carbonates. — The carbonates are the salts of carbonic acid, and the carbonates of three metals rank all others in importance. They are calcium carbonate, the sodium carbonates, and potassium carbonate. Calcium carbonate ($CaCO_3$) is the builder. That thin, white line which separates the brick or stone that is the foundation of your house, those marble pillars which stand at the entrance of



MARBLE

The Taj Mahal. Calcium carbonate in its most beautiful form as marble.

public buildings, those coral reefs which form the basis of islands in southern seas — all are calcium carbonate. In nature, wherever coral builds its shell, wherever marble quarries lie gleaming to the sky, or wherever limestone forms the backbone of huge ranges of mountains, this commonest of the carbonates exists. Man has searched for many years to find some substitute for this important compound, but in vain, and the huge stores of it which exist in the earth constantly pour forth their wealth to build the habitations of man.

Limestone. — Of the various forms in which calcium carbonate is found, limestone is the most valuable. It forms the basis for Portland cement, which, molded into blocks, supports the foundation of many structures, and protects the base of lighthouses and bridges in deep water. From this carbonate comes the lime which, mixed with water and sand, forms mortar, the common mortar which is the connecting link that holds together the brick walls of every building. Despite the great utility of the building materials which come from limestone, the increasing desire for beauty and stability in public structures keeps up a constant demand for marble, the second form of calcium carbonate, while the sea coral is forever building new islands made from this same compound in the tropics.

The carbonates of sodium. Washing and baking soda. — Next in importance, however, to the carbonate of calcium comes that of sodium, common washing soda (Na_2CO_3), which is found in nature in the soda lakes, scattered far and wide, from the hot desert sands of Egypt to our own Northwest. Baking soda ($NaHCO_3$), (commonly called in the household bicarbonate of soda) is also found in such bodies of water, and they are of no small importance as storehouses for chemicals. Ever since Bible times their existence has been known, and as far back as the book of Exodus we find a passage saying that Moses and the Israelites, fleeing from Egypt, came to a lake of bitter water which they called Marah. In our own country we find a similar place in the great Salt Lake of Utah, which holds so much solid in solution that no man can sink into its depths. These reservoirs contain bromides and iodides, carbonates and sulphates, in solution, and from them have come many important compounds, though none of these have so wide-spread a commercial use as have the carbonates of sodium.

Potassium carbonate. Potash. — More potassium carbonate (K_2CO_3), commonly called potash, exists in nature than sodium carbonate, but less than calcium carbonate. Ordinary granite contains five or six per cent of potassium compounds, and in these hard rocks is found the origin of the potash which we find widely spread through plant life. Our grandmothers well knew that potash came from wood ashes, they knew that if boiling water was poured over the dead embers of their wood fires a solution resulted which, when boiled down, gave lye or potash, which they could use for soap. But they never

POTASH

The granite rock in the forest yields potash as one of the products of its disintegration.

dreamed that the granite rocks about them furnished the wood with that substance. Yet so it was; for the rocks, constantly breaking down under the action of summer rain and winter snow, gave their portion of potash to the ground, whence tree and shrub took it as a part of their food. For many years the world's supply of potash came chiefly from the leaching of wood ashes, from their washing by means of percolating water. To-day some potash still comes from that source, but more is derived from the refuse left after beet-sugar making, from sheep's wool, and from the great natural deposits of potassium salts such as those near Stassfurt in Prussia.

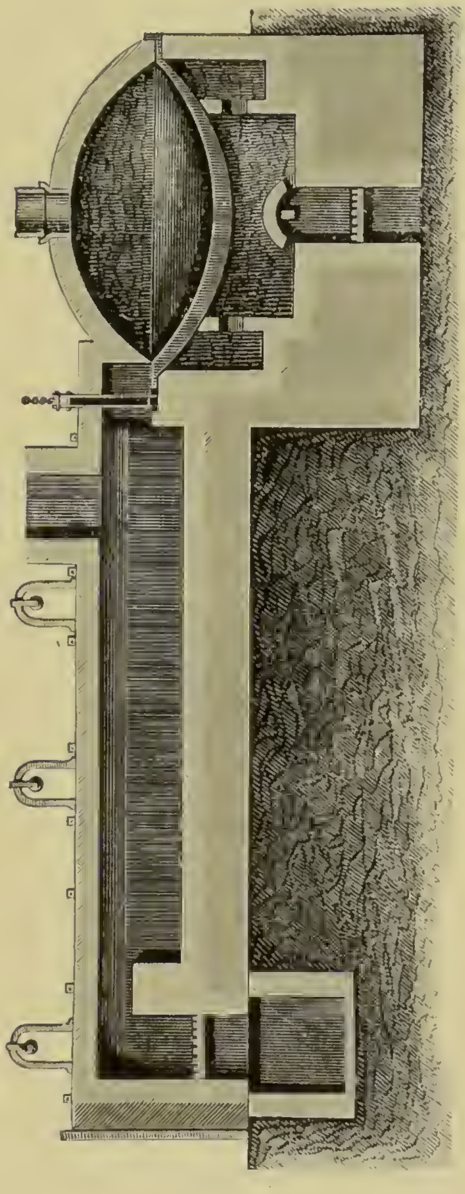
Tests for the carbonates. Their formation. — No matter under what form these carbonates are found, they always have the same chemical composition, they always respond to the same test and give off the gas carbon dioxide when attacked by an acid. The reaction for testing a carbonate (sodium carbonate) is $Na_2CO_3 + 2HCl = 2NaCl + H_2O + CO_2$. On the other hand, many metallic carbonates may be formed by running carbon dioxide gas into a solution of an oxide or hydroxide of the metal, the same action which in forming calcium carbonate has already furnished us with a test for carbon dioxide gas.

Processes for preparing sodium carbonate. — This method, though available for the laboratory, would not produce anything like the quantity of sodium carbonates required for commerce and, in consequence, this salt is obtained from one of two processes, either the Le Blanc or the Solvay. The manufacture of calcium carbonate is rendered unnecessary by the great stores of it which exist in the earth. Potassium carbonate is made either by a process similar to that used in producing sodium carbonate or by extraction from wood ashes.

The Le Blanc process. — The processes for obtaining sodium carbonate illustrate admirably the way in which commercial processes develop.

Back in the eighteenth century, when Louis the Sixteenth still ruled in France and the Colonies of the British Crown in

America were just beginning to stir uneasily under taxation without representation, soap-makers and glass-makers were using for their wares a sodium compound which came from Spain. This was prepared by burning a sea plant, known as salsoda soda, and the ashes of this plant, known as barilla, yielded about twenty-five per cent of sodium carbonate.



Le Blanc Process

The reverberatory furnace of a Le Blanc process plant. In this furnace the salt cake, mixed with limestone and coal dust, is heated on the hearths by hot gases which, rising from the fire at the left, play down from above. The heated gases pass on over large evaporating pans.

Then followed the upheaval of the French Revolution, the wars of the first Republic, and the advent of Napoleon. War had raised prices. Spain was hostile. Sodium carbonate could scarcely be procured. Napoleon determined to get some means of replacing barilla, some way of encouraging manufacture at home. He offered in consequence a prize for the discovery of a practical process for making the carbonate of soda. A French chemist, Le Blanc, brought forward a process by which it could be manufactured from common salt, and in fact invented so good a process that it is in use to-day. Briefly, this is it. Salt and sulphuric acid are mixed together. Hydrochloric acid is given off and sodium sulphate left behind, as in the action which took place when we made hydrochloric acid. The hydrochloric acid, set free, is run into water, and sold as a by-product, and the sodium sulphate is heated in a reverberatory furnace, where the flames rising from the fuel are beaten down upon the substance below, being reflected from the furnace's roof. The substance left after heating, known as salt cake, is then ready for the second stage of the process. In this the salt cake is broken up, mixed with limestone and small coal, and again heated in the reverberatory furnace. Carbon monoxide is given off and sodium carbonate, lime, and calcium sulphide are left behind. From this mixture an impure sodium carbonate, called soda ash, is dissolved out by water. To get pure sodium carbonate the soda ash is again heated with coal and the other soda compounds are changed by the carbon into a carbonate form.

The reactions for the Le Blanc process follow:

1. $NaCl + H_2SO_4 = NaHSO_4 + HCl$.
2. $NaCl + NaHSO_4 = Na_2SO_4 + HCl$.
3. $Na_2SO_4 + CaCO_3 + 2C = Na_2CO_3 + CaS + 2CO_2$.

This third reaction may be divided into two others which go on simultaneously:

$$Na_2SO_4 + 2C = Na_2S + 2CO_2.$$
$$Na_2S + CaCO_3 = CaS + Na_2CO_3.$$

Advantages of the Le Blanc process. — The great advantage of the Le Blanc process is due, first, to the facts, that the four raw materials, salt, coal, limestone, and sulphuric acid, are all common and inexpensive, and that the furnace and plant can be put up at fairly low price; and second, to the value of the by-product formed. Had it not been for the hydrochloric acid produced by the Le Blanc process it is entirely possible that other methods of manufacture would have replaced it, but the by-product here, as in many places, has saved the day and the original process is still in use.

The Solvay process. — In the other general process, the Solvay, a salt solution is saturated first with ammonia, and then with carbon dioxide. Sodium acid carbonate containing one atom of hydrogen is formed, and ammonium chloride. The acid carbonate of sodium has been so long called the bicarbonate of soda that too often the idea persists that it possesses two carbonate radicals. It does not. It is made up of one atom of sodium, one of hydrogen, and the carbonate radical. This bicarbonate in the Solvay process separates out, is heated and changes to sodium carbonate, carbon dioxide, and water. The reactions for the Solvay process follow:

$$H_2O + NaCl + NH_3 + CO_2 = NaHCO_3 + NH_4Cl.$$
$$2NaHCO_3(\text{heated}) = Na_2CO_3 + CO_2 + H_2O.$$

Solubility of the carbonates. — The solution properties of the carbonates vary considerably. Calcium carbonate is practically insoluble. Coral islands and marble pillars have stood for centuries unchanged by the washing action of both fresh and salt water. Potassium carbonate and sodium carbonate, on the other hand, depend largely on their solubility for the processes by which they are manufactured. All are, however, similar in appearance when crushed, being white powders. All are of medium density. All have a slightly alkaline reaction and all are tested, as we have said, by means of their evolution of carbon dioxide when attacked by an acid.

Carbonic acid. — Carbonic acid, from which these salts would appear to be derived, is an acid concerning whose very existence there is more or less question. When carbon dioxide is run into water a solution is obtained which turns blue litmus to a wine-red color. Under such circumstances, an acid, carbonic acid, may be supposed to be produced by adding water to this non-metallic oxide, the hydrogen of the water becoming the essential hydrogen of the acid. This substance is, however, at best, of but little practical value, though it may be considered as the basis of the great group of carbonates.

Uses of carbonates. — The uses of these bodies in nature vary, as we have seen, from the building of a spiralled shell by calcium carbonate, to the feeding of tree and shrub by the potassium carbonate of the soil. Their service in spring waters used for various medicinal purposes, their appearance as ores, as in the case of copper and iron carbonates, might also be mentioned; but all these uses are slight in comparison with the enormous commercial use of the carbonate of calcium as a basis for lime, and of the carbonates of sodium and potassium for glass-making, for the manufacture of cloths, for cleansing and for other manufacturing processes.

Mortar. — The service which limestone (impure calcium carbonate) performs in yielding lime (calcium oxide) when heated can scarcely be too much insisted upon. Wherever new buildings are in process of erection, the passer-by sees clouds of steam rising from the trough of mortar which stands before the rising frame. In mixing mortar unslaked lime (calcium oxide) (CaO), is first slaked by the addition of water, and slaked lime (calcium hydroxide) ($Ca(OH)_2$), is formed. $CaO + H_2O = Ca(OH)_2$. This is then mixed with sand (silicon dioxide). The soluble calcium hydroxide of the mixture has the property of taking carbon dioxide from the air and making an insoluble calcium carbonate, which forms a solid layer, cementing the stone of a wall or the bricks of a chimney together. This is called the setting of mortar. The chief use of the sand in mortar is to prevent undue shrinking as the mortar dries.

The use of limestone in reducing iron. — Limestone has many other uses besides that of yielding lime. It is mixed directly with iron ores to reduce the iron to its melted state, and it is an essential agent in many other processes.

The use of sodium carbonate in glass-making. — The use of sodium carbonate in glass-making is still one of the factors of modern trade, and despite the inroad of improvement is likely to remain so. Window glass, in its simplest form, is made by fusing together sodium carbonate and sand, silicon dioxide, with the addition of a small amount of some metallic oxide to increase its resistant powers. The transparency of glass is the only quality which we ordinarily consider, but its permanent quality and the ease with which it can be made to assume any form when fused are of great value in making its use general.

GLASS MANUFACTURE

A glass works with its huge chimneys rising from the furnaces below.

Window, plate, and crown glass. — Window glass employs calcium carbonate in the form of chalk as well as sodium carbonate. Plate glass does the same but adds a potassium compound. Crown glass for eyeglasses and spectacles must contain no green tint, and is therefore made with potassium instead of sodium carbonate, as is also the Bohemian glass, such as is used in beakers.

Chalk. — One thing more must be considered before closing this chapter, the use of chalk in the schoolroom. Calcium carbonate in the form of chalk leaves, as every schoolboy knows, a line of white dust against the blackboard. While various substitutes for board work have been introduced in recent years, thousands and hundreds of thousands of children have received more instruction from white chalk and a blackboard than from anything else, save the black lines left by the lead pencil on the white surface of paper.

CHAPTER XIII

EQUATIONS

Chemical affinity. — When dark iron combines with the gases of the atmosphere and leaves red iron-rust behind, when silver nitrate, added to sodium chloride, produces a white soft cloud of silver chloride, which collects in magic foam throughout a beaker, when a thousand colors, a thousand forms, are produced by the combination of elements and compounds, the thoughtful mind must seek a cause for such effects. Why does the iron unite with the oxygen of the air rather than with the more plentiful nitrogen? Why does the silver leave the nitrate to which it was joined and become bound to chlorine forming silver chloride? What principles, what laws, govern the combining of substances so diverse as the carbon dioxide of the air and cinnabar, the sulphide of mercury found in the depths of quicksilver mines? To many such a question we are forced to answer — we do not know. To some we can at least give the partial answer, — the action is due to chemical affinity. Chemical affinity may be defined as the attraction which one chemical body possesses for another. The fundamental reasons for such attraction are shrouded in a darkness as complete as that which surrounds the causation of electricity. We are to bring forward here in the following chapter some of the simpler rules which seem to some degree to explain the mysterious attractive powers which elements and compounds appear to possess for each other, for in this attraction lies the basis of chemical affinity. Together with our consideration of that force we shall take up the possibility of writing reactions or equations which shall represent in symbols the actions which take place.

A workman who can handle a molecule. — One of the great

English physicists, a man of brilliant imagination, once desired to obtain a means for readily handling such almost infinitely small particles as atoms and molecules in the work rooms of the brain. To do this he conceived of a Dæmon, a supernatural creature, who could take a drop of water and select a single molecule from the midst of its billions on billions of associates. The molecule selected from the throng, the physicist assumed that this supernatural being could carry it off to his laboratory, work with it at his bench, handle it as he chose, combine it with any other single molecule. To gain understanding of the reactions which take place in the realm of the infinitely minute, we can scarcely do better than to borrow the idea of such a Dæmon as had the English physicist, and use molecules and atoms as if we could readily see and handle them. Before taking up the meaning of the chemical symbols which we use in expressing the relations of the atoms and molecules, it may be well to consider briefly the way in which these symbols procured their names.

The symbols of chemistry. — The long corridors of the British Museum, filled with the wealth of ancient civilization, contain few memorials of the past more fascinating than the Rosetta stone, on which Egyptian hieroglyphics and Greek characters written side by side once served as a key to unlock a whole literature long hidden from the world. The hieroglyphics of the alphabet of chemistry are filled with scarcely less historic value than were the books of the pyramids translated by the stone inscription. Every symbol is full of meaning, and every element has a symbol which represents a single atom of that element. The letter O, which stands for an atom of oxygen, means not only one atom of oxygen, but also its relative weight referred to a fixed standard. The letter H stands for the weight of an atom of hydrogen. So George Jones, taken as a symbol of a man, might stand for a man weighing 150 pounds, and with a volume of between two and three cubic feet. As has been stated, the weight of one atom of hydrogen has been taken as a standard of all atomic weights, and has been given a name. It is called a microcrith. We shall use this standard for atomic weights as we use the gram for common weights. The letter O according to atomic measures, when it stands for oxygen, represents 16 microcriths of weight, is sixteen times as heavy as the atom of hydrogen. H stands as the symbol of hydrogen, and H written in a reaction means that there takes part in the chemical change which occurs one atom of hydrogen which weighs one microcrith. An atomic symbol represents one atom. Many of the elements contain two atoms to the molecule so that their molecular formulæ read with the atomic symbols doubled. H_2 stands for the molecule of hydrogen which contains two atoms. The molecule of hydrogen weighs two microcriths where the atom weighs one. O_2 stands for the molecule of oxygen which contains two atoms. O_3 on the other hand represents the molecule of ozone which contains three atoms. The molecule of oxygen weighs 32 microcriths, while the molecule of ozone weighs 48, two and three times the weight of oxygen respectively. All compounds of course contain more than one atom. Their formulæ represent molecules. In water we have H_2O as the formula of the compound molecule.

THE ROSETTA STONE

It is now in the British Museum. The picture shows only a portion.

In this compound, since there are two atoms of hydrogen weighing 2 microcriths and one of oxygen weighing 16 microcriths, it is evident that if the Dæmon we conceived put his molecule of water on a microcrithic scale he would have a weight of 18 microcriths. Suppose now that we turn to sulphur dioxide gas. Sulphur has the symbol S. Its atom weighs 32 microcriths. One of its oxides, sulphur dioxide, has the formula SO_2. That symbol tells us that there is one part of sulphur to two of oxygen in the molecule, and that the weight of the sulphur atom, 32 microcriths, added to the weight of the two oxygen atoms, 32 (16×2), makes a total weight for the molecule of sulphur dioxide of 64 microcriths.

Combination by weight. — These calculations hold good no matter what amounts of the elements combine. Two parts of hydrogen unite with sixteen parts of water whether we measure in microcriths, in grams or in pounds. Thirty-two parts of sulphur unite with thirty-two parts of oxygen to form sixty-four parts of sulphur dioxide gas no matter what standard weight we choose. To turn this to another form, suppose we assume microcriths as the standard and ask the question, What do we mean when we write $H_2 + Cl_2 = 2HCl$? Briefly, we mean that a molecule of hydrogen weighing 2 microcriths united with a molecule of chlorine weighing 71 microcriths to form two molecules of hydrogen chloride gas weighing 73 microcriths.

Where some of the elements obtain their symbols. — The symbols of most of the elements are taken from the first letter of their names; N is the symbol for nitrogen, C for carbon, I for iodine, Pt for platinum, and so on. The full list may be found in the table on the cover page. Such naming is essentially modern. When we come to the symbols of the metals known in ancient times, there blows towards us a breeze straight from the heart of the past, from Roman splendor and Grecian love of beauty. Na, symbol for sodium, comes from Natrium, a name which takes us back to the old salt works near Syracuse where "the Mediterranean floated up her waves and filled the wastes with sound." Au, symbol for gold, stands to-day for that element's atomic weight and character, but stands as well for Aurum, the gold of the Roman eagles, those golden birds which pierced through northern snows and haze of tropic heat as they went steadily forward ever widening the mighty zone of the empire. Sn comes from Latin Stannum, say the dictionaries. Fill in the color that surrounds the word, and the great sweep of Roman civilization through Britain depends upon it, for it was the need of tin, of stannum, that sent Roman legions to conquer the isle which held the tin mines.

Whence other elements derived their names. — So we might go on. Hydrargyrum gave mercury its symbol Hg. Trippingly runs the Homeric phrase, ὕδωρ ἄργυρος (the silver water) which gave the element its name for the qualities which were discovered when the swift errant metal first was found. Fe is ferrum, symbol for iron, the metal of ancient sword and shield. Pb, plumbum, is the symbol for lead, the lead that weighed down, hundreds of years ago, the nets of the fishers on the southern shores of Italy. Lead of like pattern weights the nets of fishermen there to-day. Cu is the symbol for copper; Cuprum, the word from which it came, recalls the age of bronze, the age of the first malleable metal, characteristic of the period that preceded the coming of harsh, resistant, clangorous iron.

The meaning of a compound formula. — Each and every elementary symbol of the list is full of meaning. Each stands for a certain weight, each compound symbol, the formula, gives the composition of the substance. H_2SO_4, formula for sulphuric acid, tells us directly that this acid contains two atoms of hydrogen, one of sulphur, and four of oxygen. It tells us also that its molecular weight is $H_2 = 2$, $S = 32$, $O_4 = 64$, $2 + 32 + 64 = 98$ microcriths. The molecule of sulphuric acid is ninety-eight times as heavy as the atom of hydrogen. HCl, formula for hydrochloric acid, shows one atom hydrogen to one of chlorine. That is, it has one microcrith by weight of the first to thirty-five and five-tenths microcriths by weight of the second, giving a total molecular weight for hydrochloric acid of thirty-six and five-tenths microcriths. A full list of atomic weights will be found on the cover page. These relative weights of atoms and molecules have much to do with the principles of the chemical arithmetic which follow.

Reversible reactions. — We may write simple reactions or equations in two ways. The older method was to make any chemical change a true equation whose parts balanced. For instance, when calcium carbonate (marble) is heated, calcium oxide (lime) is left and carbon dioxide goes off as a gas. In symbols according to the older usage this would read (Ca being the symbol for calcium):

$$CaCO_3(\text{heated}) = CaO + CO_2.$$

In actions like these, however, where one of the products formed is not being removed from the system, there is continually going on an action contrary to the direct action. In general it is when one of the products is removed, as when a precipitate is formed and removed or a gas goes off, that the reaction approaches completion. Beyond a certain point under other circumstances there will always be subordinate action by which some calcium oxide and carbon dioxide will tend to join together and form the original calcium carbonate from which the two spring. The action goes on in both ways, and, in consequence, we often write our reactions to-day with arrows pointing in both directions, indicating that the action is not quite completed. These arrows do not mean that there is a strong movement in a reverse direction. The greater part of the calcium carbonate, to use the special example here given, will turn to calcium oxide and carbon dioxide. Yet, since under such circumstances a portion of the original substance may be re-formed by the combination of the parts, many chemists use arrows indicating incompleteness. Using this style of expression the equation will appear as follows:

$$CaCO_3 \leftrightarrows CaO + CO_2.$$

Synthetical, analytical, and metathetical equations. — Up to the present time we have chiefly met three classes of equations. These are the addition or synthetic type of equation in which two substances are added together. This type is shown in the burning of sulphur in the air ($S + O_2 = SO_2$). Second, we have found the division or analytical equation in which a substance is broken down into its parts. An action of this type occurs when oxygen is prepared by heating mercuric oxide (HgO (heated) = Hg + O). Third, we have found the exchange or metathetical equation where two or more substances in a reaction exchange their component parts. Such a reaction takes place when we add silver nitrate to sodium chloride. ($AgNO_3 + NaCl = AgCl + NaNO_3$.) Discussion of more complicated reactions would be beyond the scope of this book so we may confine ourselves here to the study of these types.

Different types of reactions. — We may divide our consideration of reactions into two parts. First — the phenomena which occur when two solids or a solid and a gas interact. Second — the phenomena which occur during a reaction between substances in solution. The differences between these two actions are very marked. The development of theories which may explain the reactions between solid bodies and a gas, the explanation of such action as takes place when iron and sulphur, heated together, become iron sulphide, or when iron burned in oxygen becomes iron oxide, is still a matter of recent history. A great part of the work done in the nineteenth century in the study of reactions has been discarded, and the twentieth century sees but the beginning of new investigations of the causes of chemical combination.

Chemical affinity considered again. — Those investigations have not reached the point where the results can satisfactorily be recorded here. One thing does appear that, to take a specific example, the great attraction which carbon shows for oxygen, the readiness with which it burns, may have some relation to the huge amounts of energy that are given off when these two elements combine. Chemical affinity, the attractive force that impels one atom to join in combination with another, may be roughly measured by the amount of energy set free by such a chemical combination. That is the briefest, most general statement of the phenomena which have been observed in the reactions of solids together or of solids and gases. We shall see more of the energy which chemical reactions can develop, in the chapter which treats of the subject considered in the next paragraph.

Reactions which take place between substances in solution. — When we come to the action between substances in solution in water we find a much clearer field, a better opportunity for explanation than we are afforded in the other cases. Fortunately most of the reactions we commonly meet take place in solution. Every gas we have studied is produced either in solution or from bodies in a state of melting or fusion, a condition which is analogous to solution. Almost every metallic reaction has taken place in solution. Every neutralization experiment took place in solution. In fact, wherever we wish to bring about chemical action, it seems as if in nine cases out of ten our first act is to form a solution. There are two kinds of solutions, those which carry electricity and those which do not. The great majority of the cases we are to consider belong in this latter class, and it is the only one we shall discuss here.

Ions, positive and negative. — When a substance like common salt (sodium chloride) is dissolved in water it breaks up to a considerable degree into positive and negative parts called ions. The way in which it breaks up will be considered at length in the chapter on the theory of solutions. These positive and negative parts together when not in solution form molecules made up of atoms, and an ion is an electrically charged atom or radical. Now these ions show two marked peculiarities. The positive ions, the atoms and radicals charged with positive electricity, refuse to unite with other positive ions. The negative ions, those charged with negative electricity, refuse to unite with other negative ions. Positive ions will unite only with negative; negative only with positive. Bodies charged with like electricity repel. Bodies charged with unlike electricity attract. Let us see what effect this theory, called the theory of electrolytic dissociation, is going to have on our equations.

The ionic condition of two reacting substances. — Suppose we take a solution of sodium chloride. In such a solution, beside some molecular sodium chloride, we have positively charged sodium ions and negatively charged chlorine ions spinning around in the liquid. If we were to express it in symbols we might say there were Na^+ and Cl^- ions floating in the solution. Now assume we have another solution containing silver nitrate. This, beside the molecular silver nitrate, will have positively charged ions of the atom Ag^+, and negatively charged ions of the radical $(NO_3)^-$. Both the solution containing Na^+ and Cl^-, and that containing Ag^+ $(NO_3)^-$ are quietly in equilibrium. It is as if we had two small crystal bowls of water in which were floating two tiny fleets, the first fleet having half its ships black and half of crimson. The blacks of this fleet are charged with positive electricity, the crimsons with negative electricity. The second fleet has half its ships blue and positively charged, and half its ships gray negatively charged. Both fleets are covered temporarily by a thick cloud, which gives them a mantle of indivisibility. Under such conditions two parti-colored fleets rocking in the crystal bowls would be no bad illustration of silver nitrate Ag^+ $(NO_3)^-$ invisible and in solution in one beaker, and sodium chloride Na^+ Cl^- also invisible and in solution in another beaker. Suppose now the tiny fleets resting in the water of the crystal bowls suddenly undergo a cataclysm. The contents of the two bowls are poured together. What happens? The blue and the crimson, a positive and a negative, join, and their junction brings them out of the fog. The black and the gray join and remain in the fog. The fleets have changed their colors. One fleet is now blue and crimson, another black and gray. Both fleets were invisible before. One fleet, the blue and crimson, is now visible, the other is still invisible. Turn from the fleets to the solution of sodium chloride and silver nitrate. When these two solutions are poured together the positive silver unites with the negative chlorine, the positive sodium with the negative nitrate. Bodies with unlike charges of electricity have joined.

$$Ag^+(NO_3)^- + Na^+Cl^- = AgCl + Na^+(NO_3)^-.$$

Insoluble silver chloride comes out of the invisible cloak of the solution and becomes in the main a solid whose molecules cease to bear an external electric charge.

Typical reactions of electrolytes. — The reaction just given is typical of the reactions of bodies which are electrolytes, that is of the bodies which carry electricity and become ionized in water. The reason why silver united with the chlorine and left the nitrate may be found in two things, first the amount of electrolytic separation or dissociation into ions of each part, and second the degree of insolubility in water. Silver chloride formed a precipitate instead of remaining as silver nitrate chiefly because the first compound is more insoluble in water than the second. The more insoluble the compound the more likely it is to form and come out of solution. Silver nitrate is soluble; silver chloride is insoluble. Therefore silver chloride forms and precipitates.

Valence. — One principle involved here must never be forgotten. Bodies charged with like electricity repel. Bodies charged with unlike electricity attract. We shall discuss that later, but even here we can get from this statement at least a fair expectation of the way in which chemicals are going to act. Hydrogen and the metals are regularly electro-positive. The non-metallic elements and radicals are regularly electro-negative. That is the first division to guide us. We can make a second one based on what is known as valence, which, as has been said, is the combining power which a radical or an elementary atom possesses with regard to hydrogen or the equivalent of hydrogen. Suppose we try to make this a little clearer by illustration. An element one atom of which will combine with or replace one atom of hydrogen is said to have a valence of one. An element possessing this property is said to be monovalent. Sodium, which exactly replaces one atom of hydrogen in forming sodium chloride from hydrogen chloride, has a valence of one. $Na + HCl = NaCl + H$. Chlorine in hydrochloric acid (HCl) has a valence of one because it combines with hydrogen, one atom to one atom. Zinc, on the other hand, when it combines with hydrochloric acid in forming hydrogen does so by replacing two hydrogens.

$$Zn + 2HCl = ZnCl_2 + 2H.$$

As the zinc atom replaces two atoms of hydrogen, zinc has a valence of two. It is bivalent.

Monovalent, bivalent, and trivalent substances. — Iron forms two chlorides, ferrous and ferric chlorides — $FeCl_2$ and $FeCl_3$. In the first, the iron atom is bivalent since it combines with two atoms of monovalent chlorine. An atom which combines with two monovalent atoms must be bivalent since we may consider that it replaces two hydrogens. In ferric chloride iron is trivalent; it replaces three hydrogen atoms, and unites with three monovalent chlorine atoms. An element may be monovalent, bivalent, or trivalent or more. It may, as a matter of fact, possess more than one valence as was shown by iron in the example just given.

Lists of substances which are electro-positive and electro-negative when ionized. — With these two factors of ionization and valence as a basis we can arrange two lists containing the most used elements and radicals. These should greatly facilitate our writing of compounds and reactions in symbols when the action takes place between electrolytes or substances that break down into ions. While this rule is limited to that group and cannot be applied outside it is nevertheless of considerable practical value. (*See table on page* 142.)

Substances Electro-positive when Ionized.			Substances Electro-negative when Ionized.		
Name.	Symbol.	Common Valence.	Name.	Symbol.	Common Valence.
Hydrogen	H	1	Oxygen	O	2
Potassium	K	1	Sulphur	S	2
Ammonium	NH_4	1	Sulphate ion	(SO_4)	2
Sodium	Na	1	Carbonate ion	(CO_3)	2
Silver	Ag	1	Nitrate ion	(NO_3)	1
Barium	Ba	2	Hydroxyl	(OH)	1
Calcium	Ca	2	Chlorine	Cl	1
Copper	Cu	2			
Iron	Fe	2 and 3			
Lead	Pb	2			
Mercury	Hg	1 and 2			
Zinc	Zn	2			

Remembering that bodies differently charged combine, that positive unites with negative, and keeping in mind the requirements of valence, you will see that according to this scheme hydrogen will form H_2O (water), H_2S (hydrogen sulphide), $H_2(SO_4)$ (sulphuric acid), $H_2(CO_3)$ (carbonic acid), $H(NO_3)$ (nitric acid), HCl (hydrochloric acid). Even water exists to some degree dissociated as H.OH (hydrogen hydroxide).

The writing of formulae in accordance with the statement made above. — In the same manner copper which may be taken as a representative of the metals forms copper oxide, CuO, copper sulphide, CuS, copper sulphate, $Cu(SO_4)$, copper carbonate, $Cu(CO_3)$, copper nitrate, $Cu(NO_3)_2$. copper hydroxide, $Cu(OH)_2$, copper chloride, $CuCl_2$. Iron chloride forms two chlorides, $FeCl_2$ or $FeCl_3$, and iron sulphate occurs in two forms as $Fe(SO_4)$ or $Fe_2(SO_4)_3$. As a matter of convenience, it is a good plan to unite radicals in parentheses, as shown in the last two paragraphs where these groups of atoms take part in equations requiring their multiplication by two or more.

Copper nitrate, $Cu(NO_3)_2$, indicates the multiplication of the (NO_3) radical more clearly than $CuNO_{3 \cdot 2}$ can do. Bear in mind also that, in writing the formulæ of inorganic compounds, we write the symbols of the electro-positive elements, like hydrogen and the metals, first. These elements precede the electro-negative elements and radicals, and we write HCl and NaCl, never ClH or ClNa.

Examples of equation writing will be found throughout the book. Study of the many examples there given, accompanied by reference to the text, is the best method of acquiring facility in this respect.

CHAPTER XIV

NITROGEN AND AMMONIA

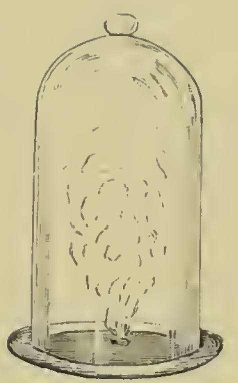
Symbol for nitrogen = N

The service of nitrogen in the atmosphere. — The gas nitrogen serves as a blanket to keep the world from burning up in the oxygen of the atmosphere. It is a restraining gas, and serves to hold back the too rapid combination of matter in oxygen. Open the draft of a stove and the resulting action comes from the increased oxygen driven through. How quickly the fire springs up in response to the added gas. Indeed, all our every-day phenomena of heat and flame come from the addition of oxygen. Suppose now that the air had a much larger percentage of this gas than the fifth part which it contains, suppose it were all oxygen. Should we attempt to live under such conditions, our bodies, like everything else which had not been completely oxidized, would swiftly burn to ashes. The slightest fire could start a great and overwhelming conflagration. A single torch might commence the destruction of the world. Fortunately, a guarding gas stands in the way, and the large percentage of nitrogen in atmospheric air serves as an obstacle to rapid oxidation. A most inert gas, it is only with the very greatest difficulty that nitrogen can be made to unite with any substance directly. The four fifths of it present in the air keeps the oxygen down to a sufficiently small proportion to enable this combining element to serve all purposes necessary to mankind, without its obtaining such mastery as would tend to wide-spread destruction.

Nitrogen in combination. — This great service of nitrogen in safeguarding the air is but a part of its worth. While it directly combines in but a few rare instances, as when it unites with red-hot metal magnesium (Magnesium nitride = Mg_3N_2), indirectly combined it is found in thousands of compounds scattered through the broad world of organic products.

One most important principle of our foods, for instance, is known by the term nitrogenous, the nitrogen containers. The proteids, those foods which are the body builders, are all nitrogenous or contain nitrogen in their composition. Nitrogen appears in combination commonly also in ammonia, in the nitrates and in the nitrites, three compounds which must be considered later.

The preparation of nitrogen. — The preparation of nitrogen in the laboratory may be effected in three ways. It may be prepared from the air by burning out the oxygen, with phosphorus, and collecting the nitrogen left. It may be made from ammonium nitrite (NH_4NO_2 (heated) $= 2H_2O + N_2$) directly or from ammonium chloride and sodium nitrite, two compounds which together produce ammonium nitrite ($NH_4Cl + NaNO_2 = NaCl + 2H_2O + N_2$). This last-named compound contains nitrogen, and yields it up quite readily when heated.



PREPARING NITROGEN

Phosphorus burning to phosphoric acid burns out the oxygen of the air and leaves nitrogen.

Properties of nitrogen. — Nitrogen is slightly lighter than air, but is so nearly of its density that it cannot be collected by upward displacement. It is neither acid nor alkali, instantly puts out a flame, does not support combustion and is chiefly characterized by the great inertness or reluctance with which it acts. The formation of magnesium nitride is one of the tests for its presence, though the negative test, that it does practically nothing that other gases do, is commonly used.

Early knowledge of ammonia. — Of first importance among the compounds of nitrogen ranks a substance known in liquid solution to the alchemists of the Middle Ages, and one whose gaseous forms were studied by that same Priestley who gave us our modern conception of oxygen. Priestley separated from ammonia water (ammonium hydroxide = NH_4OH) the gaseous substance, ammonia gas (NH_3), and named it alkaline air, for the reason that it was the only gas then known which could give an alkaline test to litmus paper.

How ammonia is found in nature. — When you read the old novels of the early part of the reign of Victoria, you find given as the proper remedy for fainting fits a vigorous inhalation of Spirits of Hartshorn. That was no uncommon compound, for Spirits of Hartshorn is the old name for ammonia. It is by no means certain that ammonia exists in its own form in the hoofs and horns of animals, but surely its component parts exist there in some sort of combination as nitrogen and hydrogen, and, when these animal substances are distilled, the nitrogen joins the hydrogen, and forms gaseous ammonia (nitrogen hydride). This gas is a common product of decay, is formed when animal or vegetable matter rots or decomposes, and occurs in rain-water, but in such slight quantities that it is almost impossible to determine its presence. When chemical tests show its presence in water of stream or lake, such tests throw out a danger signal, for drinking water which shows ammonia is only too likely to have been polluted by sewage. In consequence any drinking water in which ammonia or nitrites are found has long been held a subject for suspicion.

Combined nitrogen in organic life. — We said in discussing nitrogen that a large part of our foodstuffs contain this element. It is an interesting provision of nature that the great stores of nitrogen in the air are of but little use for the feeding of mankind since uncombined nitrogen has no power of entering into combination with most of the vegetable world. To nourish organic life we must use combined nitrogen. Ammonia and the nitrates are the forms which chiefly supply this element to the nitrogenous plants, and many fertilizers, as we shall see, depend largely for their efficiency upon the ammonia present. Decaying manure heaps almost always give an ammonia smell. Guano and other fertilizing products, including ammonium sulphate, are used largely for enriching the soil, the latter substance being used especially to fertilize the sugar beet.

The preparation of ammonia gas. — Like carbon dioxide, ammonia is formed in the laboratories of nature. The fumaroles of Tuscany are small volcanic geysers not unlike the geysers of Yellowstone Park, in which ammonia is found in company with boric acid, and from which constant jets of steam containing ammonia as gas and in solution come forth together with boric acid. In the chemical laboratory ammonia gas can be prepared through the combination of nitrogen and hydrogen by means of the electrical spark. This method is rarely used, as the reaction soon stops of its own accord, unless the ammonia formed is rapidly absorbed by an acid. A far commoner method of preparation is found in the heating of ammonia water (ammonium hydroxide). This substance is composed of ammonia gas and water, and when heated breaks down into its constituent parts, the gas going off. NH_4OH (heated) $= NH_3 + H_2O$. Our ordinary household ammonia is largely derived from the ammoniacal liquor of the gas works, a liquid left after coal has given off its gas.

Tests for ammonia. — The common tests for ammonia are its peculiar pungent odor, wholly unlike the smell from any other gas, and its alkaline action which turns red litmus paper blue. When ammonia is in combination, as in ammonium chloride, for example, a common test for its presence is the gaseous ammonia produced when the compound is heated in combination with lime (calcium oxide). This gas is recognized by its smell and alkaline action. Test for ammonia in an ammonium salt (ammonium chloride), $2NH_4Cl + Ca(OH)_2 = CaCl_2 + 2H_2O + 2NH_3$.

Cold storage by the use of ammonia. — If you have ever stood in a big market where a cold storage system cooled the refrigerator, you have seen long pipes, crusted outside with ice, which the market man said, "conducted cold." Such cold storage systems depend largely on ammonia, and similar plants are used for making artificial ice. Of all the gases, ammonia is probably the most practicable for cooling purposes on account of the ease with which it is liquefied, on account of the readiness with which it changes from the liquid to the gaseous state, and on account of the large amount of heat which is required to change liquid ammonia to gaseous ammonia, a quantity of heat which is taken from everything surrounding the ammonia which is undergoing the change. At only 40 degrees F. gaseous ammonia changes to a clear liquid, while a comparatively slight increase of pressure will make it liquefy even above the freezing-point.

Artificial ice. — These two characteristics, the ease of liquefaction and the heat required to change the liquid back into a gas, make ammonia of practical use for producing cold. From them has come the great blessing of ice throughout the tropics, where ice machines now furnish many cities with artificial ice. Not only in the south, but in many places in the north, artificial ice, produced by ammonia processes, is used to prevent the irregularity of supply caused by unusually high temperature in winter, and the resulting diminution of the ice crop.

In artificial ice machines ammonia gas is generated by boiling a concentrated solution of ammonia water in an iron boiler. The pressure of the gas evolved soon becomes so great as to liquefy a part of it. The liquid formed is run into pipes, which enter a tank containing brine. There the ammonia is suddenly released from pressure and rushes out in the form of gas. Accompanying this change comes a great removal of heat from everything around, and the water, placed in large cans within the brine solution and losing heat like everything else, turns to ice.

The use of brine in cooling plants. — In cold storage plants the ammonia gas flows through pipes surrounded by brine. The brine is reduced to a low temperature by this means and is pumped through pipes to serve as a cooler for storage warehouses, to be sent through refrigerators in markets, or into any room where a low temperature of the air is desired.

Cleansing with ammonia. — Ammonia, either as ammonia water or as the gas itself, is a great cleansing agent, removing grease and acid stains, cleansing dirt and filth. It is a valuable aid to dyeing and performs many other services to commerce.

Ammonia gas and ammonia water. — Ammonia gas must be carefully distinguished from the common household ammonia. This latter substance, strictly speaking, is ammonium hydroxide or ammonia water, and is produced when ammonia gas run into water combines with it. This liquid strongly resembles sodium and potassium hydroxides, and the radical ammonium (a group of atoms which contains a nitrogen atom and four hydrogen atoms) (the radical ammonium may be considered to have the formula NH_4) forms a whole series of compounds, nitrates (ammonium nitrate = NH_4NO_3), chlorides (ammonium chloride = NH_4Cl), sulphates (ammonium sulphate = $(NH_4)_2SO_4$), carbonates and the like, which are quite similar to the compounds formed by sodium and potassium. In consequence of this similarity many attempts have been made to separate this group of atoms which acts so much like a metal from its compounds. All such attempts, however, have been in vain, as the combination has invariably broken down without separating out as an individual substance. The hypothesis that a compound metal ammonium might exist has not yet wholly disappeared, and references are sometimes made to the theoretical metal ammonium.

The five oxides of nitrogen. — The five oxides of nitrogen, that remarkable series which, as we have already seen, is of special value in illustrating the law of multiple proportions, deserves a word or two here in connection with a brief review of the properties of the series. Nitrogen has five oxides whose symbols may give some clearer conception of their structure.

Nitrous oxide, N_2O.
Nitric oxide NO, $2NO$ (following the law).
Nitrogen trioxide N_2O_3.
Nitrogen peroxide NO_2, N_2O_4.
Nitrogen pentoxide N_2O_5.

Nitrous oxide. — Nitrous oxide is formed by the decomposition of ammonium *nitrate*, which, when heated, breaks down into water and nitrous oxide gas. This reaction should be carefully differentiated from the production of nitrogen from ammonium *nitrite*.

Nitrous oxide as an anæsthetic. — Nitrous oxide, the laughing gas of the dentist, is a colorless heavy gas with a slight pleasant smell and a sweetish taste. It is somewhat soluble in water and decomposes readily into nitrogen and oxygen. A red-hot splinter of wood rekindles when thrust into the gas. This compound can, however, be distinguished from oxygen by its action when added to nitric oxide, the second oxide of the series. Together these gases show no color changes. Oxygen added to nitric oxide gives out red vapors.

For medical and dental use nitrous oxide is to-day being superseded by some of the organic anæsthetics. At first, when it is inhaled it produces an exhilaration which is often accompanied by involuntary laughter; shortly after this stage the gas causes a state of insensibility. If the inhalation be continued too long death ultimately results, but in small quantities this gas is a safe and satisfactory anæsthetic.

Nitric oxide. — Nitric oxide is ordinarily made by the action of nitric acid on copper, the products formed being copper nitrate, water, and nitric oxide. It is a colorless, slightly soluble gas, not easily decomposed into its elements. It is commonly produced in nitric acid reactions in which it always unites with the oxygen of the air to form red-brown nitrogen peroxide. The formation of this substance is the most characteristic test of nitric oxide.

Nitrogen trioxide. — The existence of nitrogen trioxide has been questioned more than once, but the latest investigations on the subject seem to show that it does exist. This rare compound is formed in small quantities when gaseous nitrogen peroxide and nitrous oxide are mixed together and condensed to a liquid. Nitrogen trioxide when formed is a blue liquid which easily dissociates into the two compounds from which it came.

Nitrogen peroxide. — Nitrogen peroxide may be made by the addition of oxygen to nitric oxide. It is the red-brown gas which appears in a test tube where nitric acid or a nitrate is being heated. It varies in color from orange through red to chocolate, accordingly as the temperature varies, and exists both as NO_2 and N_2O_4 at different temperatures. Nitrogen peroxide decomposes only with difficulty, is suffocating and very poisonous. It forms a colorless solid at low temperatures.

Nitrogen pentoxide. — Nitrogen pentoxide is formed by burning nitrogen in pure oxygen, or in the oxygen of the air by means of the electric spark. It is also produced by removing the water from nitric acid by means of an oxide of phosphorus, phosphorus pentoxide, which absorbs water. It is a white crystalline solid which easily decomposes and rapidly absorbs water or moisture to form nitric acid.

CHAPTER XV

NITRIC ACID

Formula for nitric acid = HNO_3

Nitric acid as a solvent. — Of all the acids used as solvents in commercial work, scarcely one is more widely employed than is nitric acid. Most of the common metals dissolve in this powerful liquid, and when it is used in union with hydrochloric acid, even the resistant metals, gold and platinum, give way. "Aqua fortis," the "strong water" of the alchemist, was nitric acid, and to it they looked for such solution of the baser metals as should enable them to throw dull lead or copper into the crucible and take out transmuted shining gold.

Occurrence of nitric acid and of its salts the nitrates. — The acid is found to some slight degree free in the air, but in nature we find its salts, the nitrates, far more commonly than the acid itself. Even in cities, these compounds may be found where badly drained houses and stables send forth their drainage into the soil. About such places the whitish brown crystals, potassium nitrate (KNO_3) or niter, occasionally appear in solid form upon the ground. The radical (NO_3) is the characteristic radical of the nitrates. Such occurrence, however, is only a small, almost an infinitesimal, part of the total natural production of these salts. Ship after ship enters our harbors laden with nitrates which have been gathered from mines and from the dry beds of old and long-forgotten seas, where the drying up of the water and the oxidation of various forms of nitrogen, especially of the nitrogenous plants and organic matter on the bottom, are believed by many to have produced the deposits of inorganic nitrates. Such deposits exist in South America and in the Orient, and have furnished for many years great storehouses of the nitrates needed for manufactures and for fertilizers.

A Saltpeter Cave
A natural deposit of niter.

The preparation of nitrates from the nitrogen of the air. — These huge natural deposits are now rapidly becoming exhausted, and there is little possibility of the finding of new and similar hoards. Thence comes the effort of the modern chemist to obtain the nitrogen of the air in such a form that it may take the place which the fixed nitrates of nature hold to-day in the enriching of the soil for the farmer, and the preparation of nitric acid for the manufacturer. Some fixation of this wide-spread gas is essential, since, as we have seen, the vegetable world can only use nitrogen in combination with other substances. Great obstacles, however, must be overcome before the atmospheric stores of this element can be made available. The ocean of air under which we live contains so much nitrogen, indeed, that the weight of it in a tiny room, 10 ft. × 10 ft. × 10 ft., would amount to about 65 pounds. Think, then, what a tremendous quantity of nitrates could be produced by passing a river of air through some apparatus which would fix its nitrogen in solid form. Such a river of air is used in a commercial preparation of nitrates in more than one modern laboratory. In one of these, about 2700 cubic feet of air is treated per minute. This great current of gas passes by night and day through a flaming electric arc, known as a disc flame, which corresponds to some degree to that arc which we see in the big light in the square. Passing through this disc flame the rushing air meets the terrific temperature of the electric flame, and a portion of its nitrogen is combined with the accompanying oxygen with production of various oxides of nitrogen, especially nitrogen pentoxide. When water is added to this we have as a result nitric acid, a reaction similar to other instances where gaseous non-metallic oxides added to water produce an acid. This same action goes on in the air whenever the lightning flash spreads from one side of the heavens to the other, and from such action comes at least a part of the nitric acid of the atmosphere. For purposes of sale, acid manufactured in this way is generally run into slaked lime (calcium hydroxide), which neutralization produces calcium nitrate, a substance easily shipped in solid form to any point required. The salts thus made have proved to be quite equal to the natural ones as enrichers of the soil.

Preparation of nitric acid from the nitrogen of the air:
Nitrogen pentoxide = N_2O_5.

$$N_2O_5 + H_2O = 2HNO_3.$$

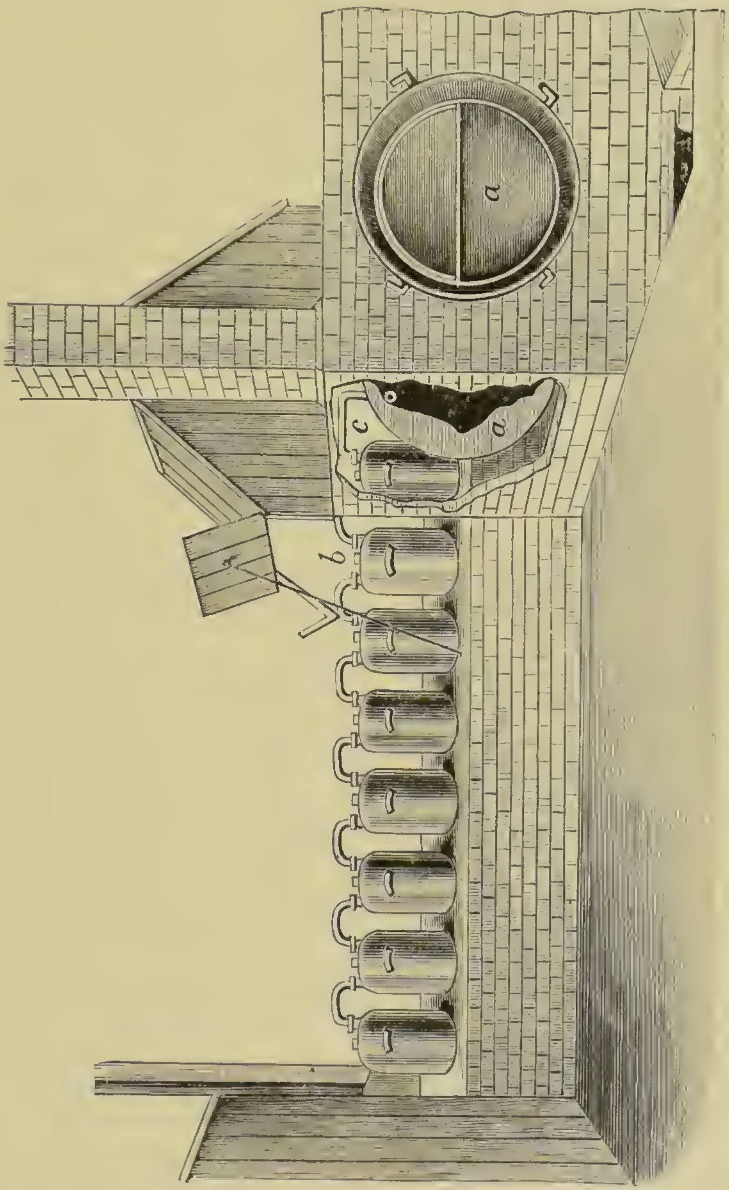
Expressing this as an addition

$$\begin{array}{l} \quad\;\; N_2O_5 \\ \underline{H_2 \quad O\quad\;\;} \\ H_2N_2O_6 = 2HNO_3 \end{array}$$

Preparation of calcium nitrate from nitric acid and lime: Lime = $Ca(OH)_2$.

$$Ca(OH)_2 + 2HNO_3 = Ca(NO_3)_2 + 2H_2O.$$

The preparation of nitric acid in the laboratory. — When manufactured by other means than the electric spark, nitric acid is commonly prepared in great cylindrical retorts lined



PREPARATION OF NITRIC ACID

Manufacture of nitric acid. The sodium nitrate and the sulphuric acid are heated in the big retort at *a*. The vapors are carried off through the earthenware pipe *c* to earthenware pots *b*.

with fire clay. In these vessels a metallic nitrate and sulphuric acid are heated together, and a sulphate and nitric acid result. If sodium nitrate, the commonest raw material, is added to hydrogen sulphate, sodium sulphate and hydrogen nitrate, nitric acid, are produced. The gaseous nitric acid which comes off from the retorts is passed through several great bottles constructed of earthenware. In these a part of the gas is condensed. From these bottles the nitric acid is generally sent through a Gay Lussac tower, a tower shaped something like a chimney, about fifty feet high and eight feet across, built of lead and filled with hard coke, through which water trickles. In this are condensed any vapors which have not been condensed in the bottles. The production of nitric acid in the laboratory is very similar to its commercial production. Sodium nitrite is commonly heated with strong sulphuric acid in a glass retort. Nitric acid comes off as a vapor. Sodium sulphate is left in the retort. The nitric acid fumes are condensed by chilling the exit tube of the retort with a wet cloth or a sponge. This experiment requires great care, *must be performed* under the hood, and is impracticable in many laboratories.

Preparation of nitric acid from a nitrate:
Sodium nitrate = $NaNO_3$.
When slightly heated this reaction takes place:

$$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3.$$

When strongly heated this reaction takes place:

$$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3.$$

Properties of nitric acid. — Nitric acid is intensely corrosive. It is colorless; its ordinary brown appearance being due to dissolved nitrogen peroxide. In all its ordinary strengths it is decidedly heavier than water. Its action on organic matter, such as textiles, has been the cause of many ruined coats and dresses in the laboratory. Hydrochloric and sulphuric acids, if left on the goods, stain cloth red, but when ammonia is added, the neutralization which occurs, takes out the stain and preserves the textile. Nitric acid stains, on the other hand, are yellow, and ammonia, while it keeps the cloth from destruction, simply intensifies the stain, making it a deeper and more permanent yellow. In consequence, great care should be taken in handling this liquid. The yellow stain which the liquid produces upon the skin is also quite permanent and can be removed in its entirety only by time, while serious burns may come from spattering drops of the acid.

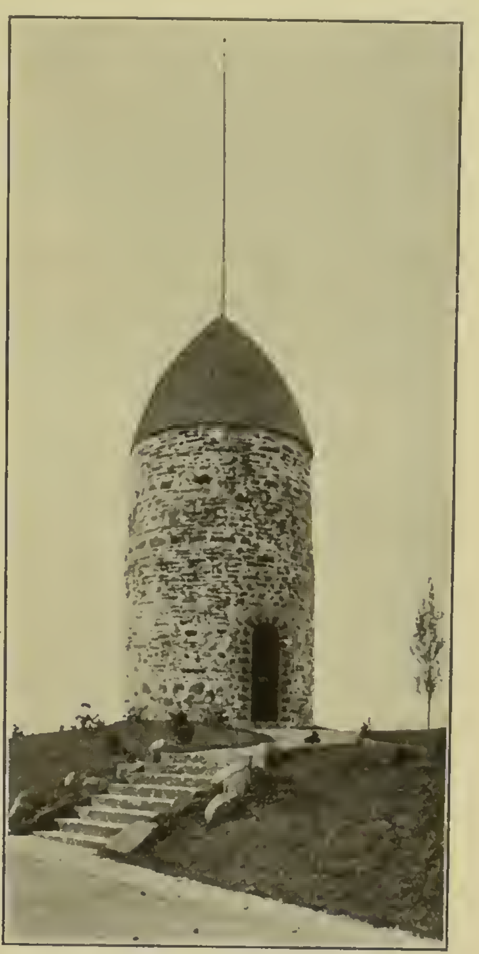
Oxidation. The oxidizing powers of nitric acid. — Nitric acid is not an especially stable compound and easily breaks down, freeing oxygen and so aiding combustion. This action gives us one characteristic apart from those possessed by the other two common acids, hydrochloric and sulphuric. When either of these act upon a metal, hydrogen gas is produced. When nitric acid is added to a metal its hydrogen, instead of being set free, joins with the oxygen of the acid and water is formed. This addition of oxygen to hydrogen is an example of a general principle, that of oxidation. The hydrogen in the case just mentioned has been oxidized to water (hydrogen oxide). Oxidation may be defined as the property of adding oxygen or its equivalent to a compound. We oxidize carbon when we burn it and form carbon dioxide ($C + O_2 = CO_2$). Reduction is the opposite of oxidation. In reducing any body we take oxygen or its equivalent away from it. For instance, when we heat mercuric oxide, the oxygen passes off and mercury is left behind (HgO (heated) $= Hg + O$). The metal has been reduced by the removal of oxygen.

The oxidation of sulphur and phosphorus by use of nitric acid. — Of the common oxidizing agents nitric acid will rank among the most powerful. If we heat sulphur with concentrated nitric acid the sulphur is oxidized to sulphur trioxide, a higher state of oxidation than if it were burned to sulphur dioxide in the air. From this sulphur trioxide, sulphuric acid is formed. $2HNO_3 + S = SO_3 + H_2O + 2NO$. $H_2O + SO_3 = H_2SO_4$. In the first of these reactions there are six atoms of oxygen in the two molecules of nitric acid. Three of these join the single atom of sulphur to form sulphur trioxide. One joins the two atoms of hydrogen to form a single molecule of water, and two join the two atoms of nitrogen to form two molecules of nitrogen monoxide. If we drop phosphorus into strong nitric acid a similar action occurs, it rapidly unites with oxygen, and phosphoric oxide is formed with the resulting formation of phosphoric acid. These two instances might be expanded into a long list.

Tests for nitric acid and nitrates. — There are two tests by which we can distinguish nitric acid: one, the blue color of the copper nitrate, formed when copper is dissolved in nitric acid; the other, the brown ring caused by nitrogen peroxide which appears when ferrous sulphate is poured carefully on to a small quantity of the acid in a tube. In this second case the nitric acid gives up a part of its oxygen to the ferrous sulphate, nitrogen peroxide is formed, and in the process a dark brown ring forms where two liquids meet. The first test would be affected by the existence of any action tending to form copper sulphate. The second would be unaffected. The test for nitrates is very similar, the only difference being that strong sulphuric acid is first added to the nitrate, when free nitric acid forms. After this the reaction goes on as given above, the blue color showing with the copper, and the brown ring with the ferrous sulphate.

Nitric acid and gold. The touchstone. — The association between nitric acid and the precious metal, gold, has been close. For centuries it has determined the true metal from the counterfeit, and, though in most modern goldsmith shops present-day methods of analysis are used, yet the old custom of the touchstone still remains in various places. Looking from the white glare of the Italian street into the dark recesses of the little goldsmiths' shops, one may see to-day the smith bending over a black stone, his touchstone. Before him he has his touch-needles, a group of twenty-four tiny bars of gold, each an alloy with a different percentage of gold, while beside him stands the bottle of nitric acid. To test an unknown alloy for the amount of gold present the goldsmith rubs the alloy across the touchstone, leaving a golden streak of metal. The nitric acid is added. All the alloyed metals, save the gold, dissolve, and the amount of gold left is tested by comparison with that touch-needle which gives a similar streak. That is the method which has been used by goldsmiths for hundreds of years, and from it the word "touchstone" has come to have a permanent place in our literature.

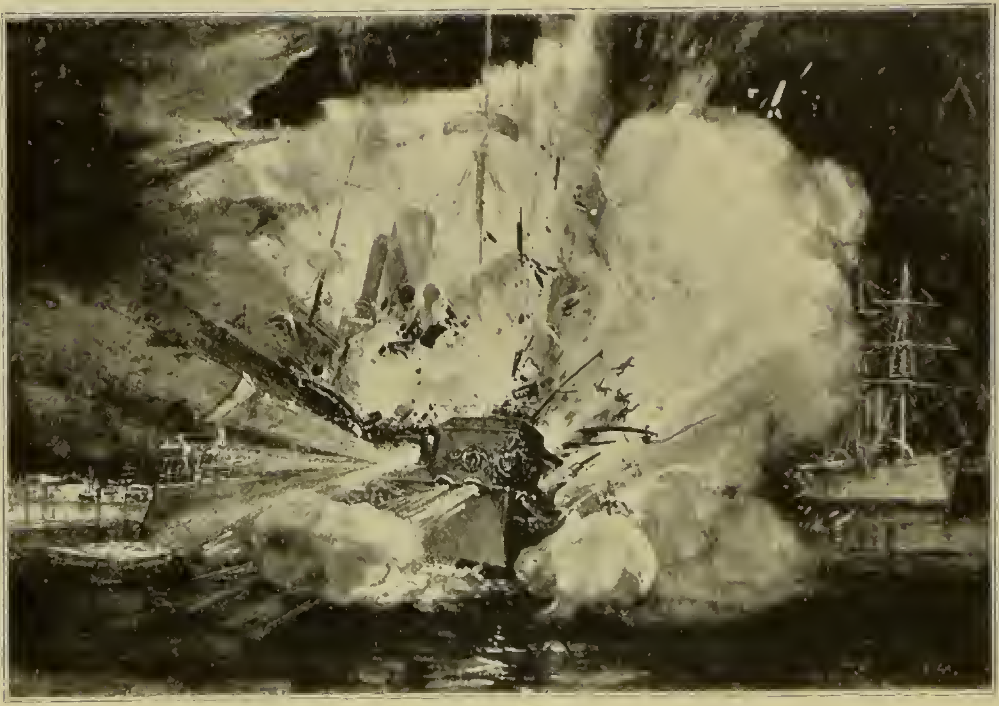
Aqua regia. — Though nitric acid alone will not dissolve the regal metal, gold, the combination of concentrated hydrochloric and nitric acids will. When these two acids join, free chloride is evolved, and gold goes into solution as gold chloride. This extremely powerful solvent has been known as "Aqua regia," the kingly liquid from the time of the alchemists to now.



An Old Powder House

Explosives in general. — Of the commercial uses of the nitrates few are more important than is their use in the making of explosives. The interest in anything which explodes, which can go off with a bang, seems to be inherent in man, and it extends from the fascination of the firecracker to the possibilities of the latest submarine torpedo. Few works of peace have called forth such strenuous efforts of research as have been shown in the endeavor to obtain explosives of constantly increasing power aimed to accomplish wide-spread destruction in case of war. For many years the expert on explosives has looked among the compounds of nitrogen for advance along his line. The production of newer and deadlier compounds of destruction by no means tends alone to warlike ends. The mining engineer must break his way into the bowels of the earth by dynamite. The civil engineer bores through great mountains by aid of high explosives. The naval engineer clears harbors, the pathways of mighty ships, by blowing up obstacles along the channel.

Gunpowder. — Nitrates are used in a large number of the common explosives. Gunpowder the oldest of them all, is the type most commonly considered to-day. Common gunpowder is a mixture of saltpeter, that is, of sodium or potassium nitrate (the "villainous saltpeter" of Shakespeare), and of charcoal and sulphur. Potassium nitrate is the salt chiefly used, as it absorbs but little moisture from the air. Sodium



Explosives I
Explosion of the U.S.S. Maine.

nitrate is, however, employed in preparing the necessary potassium nitrate. The three bodies, charcoal, sulphur, and potassium nitrate, when mixed together have absolutely no effect on each other at ordinary temperatures. When they are heated together, they react with the evolution of very large quantities of gas, the potassium nitrate oxidizing the carbon and sulphur, and producing quantities of gaseous oxides of carbon. The explosion of the firecracker and the speed of the bullet depend primarily on similar action for their effect. Indeed, we may say that explosions in general produce their effects by the sudden conversion of small quantities of solids or liquids into very large quantities of gases, which are generally greatly expanded by the heat generated during the chemical change. Such action takes place in every direc-

Explosives II
Wreck of the Maine. Details looking forward after the explosion.

tion, though it is hard to realize as we look down into a deep pit which dynamite has excavated that a far greater hole than that we see has been made in the atmosphere pushed aside by the gas resulting from the explosion. When a bullet is shot from a gun it is driven forth by the expanding gas, though under these circumstances the action is at first confined to the barrel of the gun, and a large part of the energy of the gas is consumed in driving the bullet forth.

The nitrogenous plants of the soil. — Though the necessity for providing nitrogen for the soil, and the attempts to obtain fertilizer from the nitrogen of the air, have already been discussed, yet there is a process not yet mentioned by which nitrogenous plants can be made to obtain their needed food from the air. The use of this process by man came into existence through the discovery of a remarkable group of bacteria, tiny organisms, which act as tiny manufactories, taking nitrogen from the air and turning it into the form of plant food. It has been known for a number of years that those nitrates which existed in the soil were not sufficient to provide all the nitrogen necessary for certain crops, as the nitrates on ordinary land are not normally plenty, being largely produced from the chemical change of the ammonia found in decaying vegetable matter and, in a lesser degree, from the rainfall. To obtain these nitrates the roots of the plants send out countless hair-like tubes through which they drink up solutions containing these salts. Leguminous crops, like peas and beans, contain far more nitrogen than they could draw from the soil, and do not benefit as other plants do from fertilizers containing nitrogen; such plants must have some means of obtaining nitrogen, unknown to other divisions of the plant world. Through careful investigation it was determined that the roots of many plants of this type furnished lodgment for living vegetable organisms, which had the power of transforming the nitrogen of the atmosphere into a form where it was available for food.

Plant factories which convert nitrogen. — Each plant of this type under these conditions becomes a factory which manufactures its product, the grown plant, from its raw material, the nitrogen of the atmosphere. When that was once discovered, the next step was to procure these organisms and add them to the plants which needed their aid. This was accomplished by the addition of cultivated bacteria to help the thousands of tiny workers, clustered about the plant roots. While this discovery has not been completely developed as yet, it is entirely probable that by its aid barren fields may be made fertile and plants may be increased in size as well as in productiveness. The farmer who desires to increase his yield of leguminous crops may then inoculate his seed or his soil, or perhaps both his seed and his soil, with nitrifying bacteria.

Sewage disposal by means of bacterial action on organic nitrogen. — The service of the nitrogenous bacteria as scavengers is another modern use of biological chemistry. The most modern sewage disposal plant is simply a large pile of stones on which bacterial films may form. The nitrogenous wastes of the city pour from the city sewer on to these rocks. There the bacteria seize upon the organic nitrogen, and, by the help of the oxygen of the air, turn the dangerous forms into harmless inorganic nitrates and nitrites, which flow from the base. The tiny organisms build into their own structure the harmful constituents and set free innocuous compounds.

Nitrous acid. — Far less important than nitric acid is the form which is nearest to it in composition, nitrous acid, whose salts are the nitrites. Coming down with the rain we sometimes find solutions of compounds which rarely or never exist naturally on the earth, but which have been brought into being by some strange combination in the skies. One such compound is ammonium nitrite, a salt of a perhaps non-existent acid, nitrous acid. Though this acid may not even exist, and though few of its salts are found in nature, yet many of its compounds, the nitrites, can be produced in the laboratory, and their characteristics are well known. They are commonly prepared by the removal of one oxygen atom from certain nitrates by gentle heating. Under such circumstances nitrites are left behind. The test for their presence is the evolution of brown fumes when sulphuric acid is added. They are particularly likely to occur in sewage, and the presence of any considerable amount of them in drinking water should be a red flag of danger, warning us against the use of the water. The formula for nitrous acid is HNO_2, for ammonium nitrite is NH_4NO_2.

CHAPTER XVI

AIR

Air has no symbol or formula. It is a mixture of different gases

The ocean of air. — Down at the bottom of the briny ocean, say the fairy chronicles of the Arabian Nights, stand cities and palaces where live a race who, spending their lives under the waves, eat and sleep, work and play, beneath the salt sea. Wild dream of the imagination as that seems, our daily life offers no stranger conception, for we, like the fabled people of the sea, live at the bottom of an ocean. Theirs is an ocean of water. Ours is an ocean of air.

The weight and depth of the atmosphere. — This gaseous ocean of ours, the atmosphere, is far deeper than the most cavernous depths of the sea. Far up into the blue empyrean, twenty, thirty, forty, fifty, and more miles, stretches this envelope whose outer verge is unknown, whose density varies from the comparatively thick atmosphere of the sea level to the thinnest gas we can conceive, in which the matter is separated by wide distances. We know roughly that on certain days the air seems heavy, on other days that it seems to hold a certain fine lightness, and we realize in a vague way that the air does not appear to have the comparatively constant density of wood or earth nor the very constant density of water. This ocean of ours is more changing than the sea. Its weight on summer night or day of fog varies greatly from its weight on a bracing winter day or in the clear moonlight of the hunter's moon in October. Such variations constantly occur on the same level. Change the level, ascend a lofty mountain, and the air becomes rarer and rarer, till men upon such heights gasp for breath to gain sufficient oxygen to feed the body fires adjusted for lower levels.

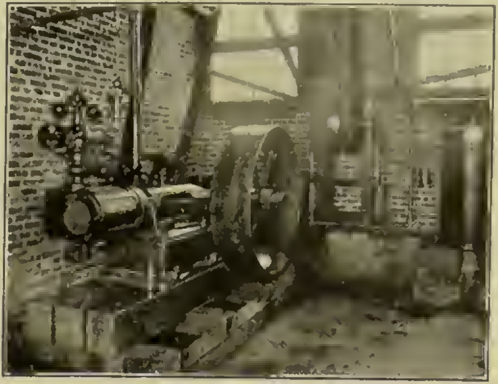
The varying weight of the air. — A balloon rising higher than the mountain's top shows a yet more striking change. The barometer, which measures air pressure as a scale measures the weight of solids, falls rapidly, the air pressure and the density of the air become constantly less, the higher we go. Let us see if we can find the explanation of the apparently irregular construction of this mantle of air which covers our solid globe as it goes journeying through space.

Atmospheric Pressure

A haycock. The varying pressures of hay from the top to the bottom of a pile may be considered analogous to the pressure of the atmosphere.

A haymow and the atmosphere. — A haymow in a country barn offers a close comparison to the atmosphere. On August days the great loads of sweet-smelling hay come entering the barn to fill the empty space. As one load is piled upon another, and then a third and fourth are added, the loose hay towards the bottom becomes more and more compact, and becomes pressed together almost like cotton in a press: the lowest layer is the densest, the next less dense, and so on until the top layer is as loosely held together as the first was when thrown upon the barn floor. The weight of the hay above compresses the hay beneath. The more the hay above, the denser the hay beneath. There is a constantly increasing compression in every haymow from top to bottom.

Compressibility of the air. — Air, like any other gas, is a far more compressible substance than hay. Whenever you blow up a bicycle or an automobile tire you are compressing comparatively large quantities of air into a small space. All the compressed-air machines, such as the drills of miners or the compressed-air automobiles, depend for their action on the extreme compressibility of air and its subsequent expansive powers. If hay increases in density because of its own weight, how much more will air compress? The air on the seashore at the bottom of this gaseous ocean is compressed most because it has the highest column of air above it. The air just above the sea level is driven together a little less, and is, in consequence, a little less dense. All the way up to the farthest edges of our atmosphere to the point where it quite disappears, leaving nothing but that unknown ether which fills the spaces between the planets, comes a constant lessening of the density due to the fact that the weight of layer after layer of air is removed as we mount upward. That is the reason for the constant lessening of air pressure, of the constant diminishing of the density of the atmosphere as the level rises. Changes on the same level are due to the shifting waves of the gaseous ocean, wherein the wind is ever sending scurrying clouds so that the pressure on the surface of the earth varies according to the variation of weight of the column above us.



Air Compression

Compressing air by the power of steam.

Composition of air. — We know that air holds oxygen, the gas that keeps ablaze the fires of all the earth, and that it holds, as well, nitrogen, the gas which puts a damper on those flames sufficient to allow them to be used safely and efficiently. We have already said that there is about one fifth oxygen and four fifths nitrogen in air. Notice that word "about," for there is left some small percentages of other things which play important parts in our daily life. Oxygen and nitrogen play the chief part, however, and make up more than ninety-eight per cent of ordinary air. It is, therefore, to their share in the life of the atmosphere that we should turn first. And the natural question rises, how do we know that there is one fifth oxygen and four fifths nitrogen in air?

Separating the nitrogen from the oxygen of the air. — Suppose we take a dry bottle filled with air, place it mouth downward over water, and, putting a bit of phosphorus beneath it in a tiny porcelain boat, set fire to the phosphorus (using great care in so doing, by the way, for phosphorus must be kept under water to prevent its burning in air and making ugly burns when handled with the fingers). In such an apparatus we should have air and a substance, phosphorus, which easily burns in air, combining with oxygen and forming an oxide. The phosphorus, once lighted, burns fiercely, combining rapidly with the oxygen of the air and depositing a white solid oxide of phosphorus on the side and top of the bottle. The phosphorus uses up all the oxygen present in the bottle while burning. It does not unite with the nitrogen. The oxide of phosphorus formed is not an insoluble gas like oxygen, which would hold the water in its place, but a soluble solid like salt or sugar. From that last fact comes our determination as to how much oxygen and nitrogen exist in air, for the moment the phosphorus oxide (phosphorus pentoxide = P_2O_5) begins to form, the water begins to dissolve it while the water beneath the bottle begins to rise, slowly but surely taking the place of the oxide which has gone into solution. Now notice two things. First, the water rises till all the oxide of phosphorus has dissolved. Second, the phosphorus used up all the oxygen in the bottle, and its oxide took the place of the oxygen. If we now see how much water has entered the bottle we can tell how much oxygen was there at first, since the nitrogen, which took no part in the action, will remain unchanged as a gas. Measuring the amount of water which has entered gives approximately the amount of oxygen which was present. Measuring the capacity of the bottle and subtracting the oxygen which was there gives us the amount of nitrogen present.

This experiment has been done thousands of times, and, properly performed, always gives the same result. Four fifths of the bottle is filled with gas, one fifth with water. Four fifths of the air is nitrogen. One fifth of the air is oxygen. The other constituents of the air, outside of water vapor, all together make up less than two per cent of the whole. Among the other important constituents are argon, carbon dioxide, and certain compounds of nitrogen. The less important are rare gases, such as helium and neon, gases which are found in the atmosphere in almost infinitesimal quantities, and which are as yet of but slight practical importance, though of considerable theoretical value.

Argon. — Oxygen has already been discussed. Nitrogen and carbon dioxide, like oxygen, play roles in the economy of nature too important to be dismissed with a paragraph: each of these gases has a chapter to itself. Argon (A), however, finds its only relation to our every-day world in the atmosphere, and therefore demands an explanation here.

Not till the close of the nineteenth century, in 1894, did any chemist even suspect that such a substance as argon existed in the air about us. During the latter half of the eighteenth and the first half of the nineteenth century many men worked anxiously and zealously to find the whole secret of the air. Lavoisier, Priestley, Scheele, Dumas, Cavendish, and Boussingault, had one by one solved question after question which had puzzled philosophers for centuries. Cavendish came nearest to its discovery, but even he passed on and left the gas unknown. For half a century the scientific world believed that all the constituents of the air were known and the tide of research had swung to a thousand avenues more directly concerned with industrial progress. Suddenly came news that shook from their pedestal the beliefs of years. Rayleigh and Ramsay, two brilliant English chemists, announced that air contained 0.94 per cent by volume, 1.3 per cent by weight, of a new and hitherto unknown element, a gas which they called argon.

The discovery of argon. — Here was an amazing example of the power of concealment possessed by certain of the natural wonders of this world. Thousands, almost hundreds of thousands of experiments had been made to determine the composition of the air, but argon had never before been found. It had concealed itself in one of the constituents of the air, in the nitrogen, so completely that whenever the nitrogen was removed from air the argon was removed with it, and was weighed and measured as a part of the nitrogen. It was only when its discoverers found that nitrogen from air weighed slightly more than an equal volume of nitrogen gained from other sources, that the possibility of finding this hidden element came. If the nitrogen of the air weighed more than nitrogen from other sources they reasoned that it must be because it contained something heavier than itself. The hidden gas was separated out and proved to be a new element slightly heavier than nitrogen, colorless, odorless, acting like nitrogen in almost every respect, but even more inert. Nitrogen will combine with a few substances. No substance has yet been found with which argon will combine.

Water vapor in the air. — The water vapor always present in the air may or may not be classed among its constituent parts according as we think of the air as dry or otherwise, but in its action it has much to do with the comfort of the individual and with many a daily task. As liquid water in a kettle on the stove soon changes to gaseous water (steam), and, by evaporation, is added to the air of the room, so the great fire of the sun is constantly evaporating the water of sea and lake, of brook and pool, drawing it up into the surrounding atmosphere. The air so charged becomes humid, and it is on the amount of water in the air that changes in humidity depend. The air at any given temperature will hold only a given amount of water vapor. When the air is saturated and that temperature is lowered, the water spills over, so to speak, and descends on the earth as snow, rain, or hail, or else coats the ground with frost or dew. In the house on a summer's day such water vapor coats the glass of ice water or the cold-water pipe. The beading of dew on cold surfaces comes from the condensation of the water vapor of the air around. On an average fair day the atmosphere is likely to hold about half as much water vapor as it can take up. On a dog-day in August the water vapor in the air may easily go above ninety per cent of the total which the air can carry. The hot, sticky feeling which characterizes such days comes from the fact that perspiration cannot readily pass into the water-charged air about us, and that much of our comfort depends on such evaporation taking place. Wet clothes on a line dry better on clear, dry days than on humid days, because on such days the air will take up the water from the clothes.

Diffusion. — The various gases of the atmosphere, oxygen, nitrogen, argon, water vapor, carbon dioxide, and the rest are each of different weight. One might naturally expect that they would rest each in layers, the heaviest at the bottom and the lightest at the top, as lighter oil rests on heavier water. No such thing happens with air. By a wise provision of nature, in any mixture of gases all the individual gases become diffused through the whole. Think how the odor of a bunch of roses permeates every corner of the whole room where they are placed. So each gas spreads through the atmosphere. All gases possess this property of diffusion. From this property comes the comparatively constant proportion of the various gases in air.

Liquid and solid air. — Air, though normally a gas, can be made liquid and even solid. A whitish liquid when carbon dioxide and ice are present, it is a beautiful pale blue when these are removed. Its temperature is hundreds of degrees below the freezing-point, and liquid air boils between three and four hundred degrees F. below zero. If liquid air is exposed to the ordinary temperature of a room it rapidly boils away and enters the atmosphere. Almost all substances dropped into this liquid become extremely hard from the great cold, while, by a strange paradox, red-hot steel burns brilliantly when plunged into it, because air in its liquid form contains generally a larger percentage of oxygen than gaseous air.

Air a mixture, not a compound. — Now for a last but very important point. One question which keeps recurring is: "How do we know that air is a mixture and not a compound?" To answer that, suppose we consider the definitions of a mixture and a compound. A mixture is made up of two or more ingredients which are wholly separate, independent, retaining all their individual properties no matter how closely they are mingled. A compound is a substance produced by a union of two or more ingredients which, when combined, form a distinct independent substance different from any of the ingredients.

Some proofs that air is a mixture. — In air every constituent acts exactly as it does out of air. The flashlight powder of the photographer is magnesium. In air it blazes up suddenly with high light and leaves behind a white soft powder which is magnesium oxide. In oxygen precisely the same effect occurs excepting that, as in all cases where substances are burned in pure oxygen, the combustion is much more energetic. As before, magnesium blazes up and soft magnesium oxide is left. This same metal furnishes a second proof. Magnesium is one of the few metals which combine with nitrogen. Its combination with nitrogen is precisely the same whether that combination takes place in the nitrogen of the air or in pure nitrogen. Whether we take carbon dioxide from the air or from some substance which holds it, we always find the actions in which it takes part followed by the same result. Returning to our definitions, then, each substance has acted in a wholly separate and independent fashion, and has retained all its individual properties. In consequence air holds very closely to the definition of a mixture, since each of its properties belongs separately to its individual parts. In every known compound the properties of the compound are different from the properties of its parts. Air has no such characteristic and cannot be a compound.

Air, then, is made up of many cunning workers. Oxygen keeps alive the fires, and keeping them gives life to man, to beast, to plant. Nitrogen keeps the fires a servant, not a master, though they unchecked would sweep the earth. Argon serves as an aid to nitrogen. Carbon dioxide acts as the food of the plants. Water vapor, fallen to the earth, gives kindly sustenance to plant-life, and so softens the air as to make life more tolerable for the animal inhabitants of the globe and for all the varied organic world that covers its crust. This wide-reaching envelope, the atmosphere, this covering of the earth, the air, touches us all, and, touching, gives more service than any other mixture.

CHAPTER XVII

THE GAS LAWS

The chimneys of our houses and their relation to the gas laws. — Night and day the gaseous wastes of the furnaces pour through the chimneys of the city from the fires below. The wavering haze which tops those chimneys in the winter is caused by the passage of light rays which come to our eye through light, hot gases that rise through the heavier air around. The draft of the fires depends on the drawing power of the chimney. The mission of chimneys is to take advantage of the laws that govern hot and cold gases, and to enable us thereby to use fire as we wish, to make of it a useful and powerful servant.

Gas Pressures

A tall chimney carrying off hot gases from the fire below.

How a chimney helps a fire. — All gases contract greatly with cold, and expand greatly with heat. No liquids or solids can compare with the gases in regularity of expansion or in the changes of volume which occur with changes of temperature. No changes of volume affect the ordinary conditions of our daily life more than do those produced by our fires. To understand these matters better let us review some simple facts. Ordinary burning, as we know already, is nothing more or less than the oxidation or combination of a fuel with the oxygen of the air, and we have seen that to have a fire we need two things, carbon or some other like substance that will oxidize, and air to supply oxygen. If we are to keep a fire burning regularly we must keep up a constant supply of fuel and a constant supply of air. In our houses the first necessity comes from the wood-pile or coal bin, and the second is supplied regularly from the room or the cellar by means of the chimney behind the stove or heater. This is a new conception of a chimney. Not only does it take off the wastes of the fire, but it pulls air over the fire. Suppose we see how that is worked.

Suction really the creation of a partial vacuum. — When you drink lemonade through a straw, you say that the lemonade rises in the straw because of suction. Suction is really the creation of a partial vacuum; that is, when you inhale your breath through the straw you pull out some of the air within it, whereupon the air inside the straw is exhausted. Under such circumstances the great weight of air in the atmosphere above the earth pushes down the lemonade in the glass and drives it up the tube. The lemonade rises in the straw because a partial vacuum has been formed within the tube and the liquid is driven up by the gases of the atmosphere outside.

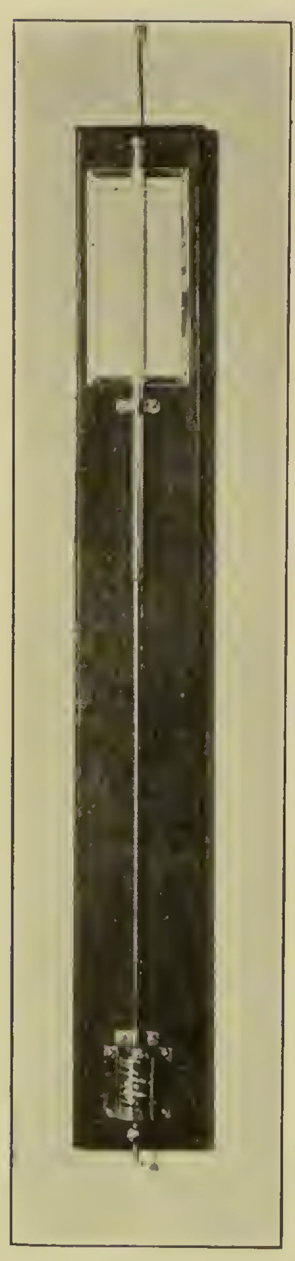
The varying density of gases. — A similar action occurs when you build a fire. Gases expand when heated. Therefore when a gas is heated it becomes lighter. To illustrate that statement suppose we take a cubic foot of a gas that weighs a pound, and heat it until it expands to double its size, becomes two cubic feet in volume instead of one. The weight has not changed but the volume, and, in consequence, the density has. Since both together weigh a pound, each of the two expanded cubic feet weighs half a pound and the density of each is half what it was before, for there would be only half as much matter for each unit of volume. Every true gas would act like this air, and we may say that any given amount of gas decreases in density (becomes lighter) the more it is heated, since the more it is heated the more it will expand.

Why hot gases rush up a chimney. — Now for the chimney. The hot gases that come off the fire and float up the chimney, although mixed with the uncombined nitrogen and air that accompanies them, weigh less than the air of the house. These gases escape into the air because their density is less than that of the air outdoors, just as a balloon which is lighter than the air which surrounds it floats upward because of its smaller density. While the fire glows the chimney is filled with gases that weigh less than the gases of the air outside. A chimney is similar to a lemonade straw from which you sucked the air. It is a tube filled with gases lighter than the atmospheric gases outside. In the preceding case the lemonade was driven up the straw by atmospheric pressure seeking to fill a partial vacuum. Here the weight of the heavier atmosphere forces fresh air through the stove over the fire and up the chimney after the lighter gases. And a large part of the reason why the gases rise up the chimney and why air rushes through the coals is because heating has expanded the gases from the air and made them light.

The law of Charles. — The changes which gases undergo with heat and cold are in accordance with a law, which is known as the law of Charles. This first of the three great gas laws, besides its advantages in every-day life, is one of the foundation stems on which our present-day theory of chemistry rests. This law says, "Under constant pressure every true gas expands or contracts by a certain definite percentage of its volume for each degree Centigrade that it is warmed or cooled." Putting it more definitely, we may say that with a fixed pressure every true gas expands or contracts $\frac{1}{273}$ of its volume at 0 degrees C. for each degree Centigrade that it rises or falls. That is a very wonderful law when one comes to think it over. "Every true gas," — think what that means. All the gases we have studied, every constituent of the air, every liquid which boils away into an invisible vapor, every breeze that blows, each and all obey this law. Each and all change $\frac{1}{273}$ of their volume for this change of one degree of heat or cold. To illustrate a bit farther, — Suppose we have a fleet of 273 ships floating on the tide in a quiet harbor, and that every hour a ship comes in. Each incoming ship is, of course, $\frac{1}{273}$ of the fleet. At the end of the first hour there would be 274, at the end of the second hour, 275, at the end of a hundred hours, 373, at the end of 273 hours, 546, a fleet double the size of the first. So with a gas. 273 cc. of a gas at 0 degrees C. increases $\frac{100}{273}$ in going up 100 degrees and becomes 373. It increases $\frac{273}{273}$ in becoming 273 degrees warmer and doubles in volume. When cooling instead of heating the same characteristics are seen. 273 cc. of gas at 0 degrees becomes 272 cc. at − 1 degree, 271 at − 2, 263 at − 10, 173 at − 100, 73 at − 200, if it goes down to that low temperature in the form of a gas. Apparently if the law held good the gas would disappear when we reach − 273, and leave a void when $\frac{273}{273}$ was subtracted from its volume at 0 degrees. Such a condition would be impossible. It would be in direct contradiction to the unchanging law of the conservation of matter. Matter cannot be destroyed. Every known gas becomes a liquid before − 273 degrees C. is reached.

Absolute zero. — That temperature of 273 degrees, however, is an extremely valuable one even if we have never been able to get a gas down to such an intensity of cold. All gases at whatever temperature they become gaseous whether at 100 degrees or 300 degrees act in expanding as if they had started from nothing at − 273 degrees, and had one unit of their volume at − 273. That is, − 273 degrees C. serves as the starting-point of all contractions and expansions of gases. It is therefore given the name "absolute zero." Absolute zero on the scale is 273 degrees below the Centigrade 0. You will find many references to this standard, and to the part it plays in computing gas volumes.

Normal or standard temperature. — The temperature of absolute zero is not serviceable for the comparison of gases, so to get a standard temperature at which we can readily compare all gases we use 0 degrees C. (the temperature at which water freezes) and we call this the normal or standard temperature. When you read of a gas being measured at N.T. or S.T., it means at normal temperature or standard temperature, which is 0 degrees C. or 273 degrees absolute. All gases are compared at one and the same temperature, and are regularly compared at 0 degrees C., for, if we used more than one temperature, the volumes of different gases expanding as they do with heat and contracting with cold would be different. You will find a numerical method for reducing gases from various temperatures to normal temperature in the study of chemical arithmetic.



Mercury Barometer

The barometer as a balance. — When you try to prophesy the weather for the afternoon from the reading of a mercury barometer, you are but little likely to realize that this long tube partly filled with shining mercury is a balance; as true a balance as the chemist's or the grocer's scales. The only difference is that the barometer balances one substance only, balancing the weight of the ocean of air which surrounds us.

How the barometer balance acts. — The action of this barometer balance is simplicity itself. Take the two opposing pans of a druggist's or a chemist's balance. If we put a pound weight on one, the other rises and the two will not balance until a pound has been placed on the empty scale pan. Equilibrium, balance, is only obtained when the weights on each side are equal. We could make a liquid balance which would act in a similar way from a piece of garden hose, or better still from a piece of glass tubing bent U-shaped. If we place enough mercury to create a certain pressure in such a tube we can balance an equal pressure of water with the mercury, for the heavy metallic liquid would rise higher and higher in its tube as the quantity of water grew more and more. By marking a scale on the tube of mercury we could tell the pressure of water which was acting upon the mercury. A fixed pressure of water would send the mercury to one height, three quarters of that pressure would send the mercury correspondingly lower. Double that pressure would drive the mercury to double the height on the scale. In a similar way we balance the weight of the atmosphere by the barometer's tube of mercury. Imagine a closed tube of glass set in a bowl of mercury. At the top of the tube is a vacuum, a space as nearly emptied of air as can be easily obtained; below the vacuum is mercury. The bottom of the tube is open and rests on the mercury of the bowl. Such an apparatus would be a simple barometer. The air presses on the mercury in the bowl, and holds up a weight of mercury in the tube, equal to the weight of air pressing on an equal surface in the bowl. If the air grows heavier the barometer column rises. If the air grows lighter the barometer column falls. The mercury rises as the atmosphere outside grows denser, falls as it grows less dense. To understand the variation in the pressure of the atmosphere and the consequent change in the barometer you can scarcely do better than review briefly the preceding chapter on the atmosphere.

The varying pressure of the atmosphere. — As we have already seen in our study of the air the pressure of the atmosphere is always changing, because of the changing pressure, and the resulting compression of the gases which compose it. It is necessary, therefore, to take careful account of this changing pressure in calculating the volume of any collected gas. The changing pressure of the air works on a liter of gas, for example, as the additional weights of a scale work when they bear down the scale pan where they rest. To-day we may have a liter of hydrogen in a confined space. To-morrow if the atmosphere is heavier we may have less than a liter in the same space. The heavy air will have compressed the gas. Day after to-morrow, if a rain is coming on, we may have more than a liter. The lighter atmosphere has allowed the gas to expand. Evidently we need some standard pressure at which all gas volumes may be compared. We obtain this standard by means of the second great law of gas volumes, the law of Boyle, sometimes called the law of Mariotte.

Robert Boyle. (1627–1691)
Discoverer of Boyle's Law.

The law of Boyle or of Mariotte. — The law of Boyle states that "at constant temperature the volume of every true gas varies inversely as the pressure that is put upon it." Varying inversely means that the greater the pressure exerted on any gas the smaller its volume will become, or, stating it conversely, the less the pressure exerted on any gas the greater its volume will become. Suppose that we have any two pressures exerted on a definite quantity of gas, we can write this law in the form of the following proportion: The first or greater pressure : the second or smaller pressure : : the second or larger volume : the first or smaller volume. Putting that proportion into terms of gas in a bottle or a reservoir which is acted on by the pressure of the air around and you may say: The pressure of heavy air : the pressure of light air : : the large volume of gas under light pressure : the small volume of gas under heavy pressure.

Normal or standard pressure. — Every gas tank that supplies gas to the city, every automobile engine where exploding gasoline turns to vapor, every pneumatic rubber tire, every balloon rising into the sky, feels changes in atmospheric pressure, and the receivers which contain the gases that they use must be so built as to allow for changes from this cause. In chemistry, because of this continual irregularity in the volume of a gas, we reduce all gas volumes to one common standard which is the normal pressure (N.P.) or standard pressure (S.P.) of gas volumes. This standard pressure is that of a column of mercury 760 mm. in height. This is a common height of the column of the barometer tube on a fair day.

Normal or standard pressure and volume. — Every gas volume is computed from the two standards set by the laws of Charles and of Boyle, and all are compared at normal temperature and pressure, that is at 0 degrees, and at a pressure equal to 760 mm. of mercury. The letters N. T. P. or S. T. P. stand for Normal or Standard Temperature and Pressure.

The law of Gay Lussac. — When, by application of the laws of Charles and Boyle, we compare like volumes of gases, which have entered into combination, we find a third great gas law, the law of Gay Lussac. This states that gases combine in volumes which bear a simple ratio to each other and to the volumes of the product obtained. To illustrate:

2 vols. hydrogen + 1 vol. oxygen = 2 vols. gaseous water.

1 vol. hydrogen + 1 vol. chlorine = 2 vols. hydrogen chloride gas.

3 vols. hydrogen + 1 vol. nitrogen = 2 vols. ammonia gas.

Three and two, two and two, four and two, all these are simple ratios, 3:2 – 2:2 – 4:2. All these ratios are made up of small whole numbers. There are no fractions. Gases never combine each with each in fractional proportions, halves, thirds, or quarters.

From these laws was evolved a great hypothesis, a brilliant theory which has been of much service to science, yet one which, as it is based on the atomic conception, is now and probably will remain hypothesis rather than a law, the hypothesis of Avogadro.

Avogadro's hypothesis. — It was about a century ago that the Italian scientist, Avogadro, proposed this wonderfully simple yet daring theory. Briefly stated it is as follows: "The same number of molecules may be found in equal volumes of all gases which have the same temperature and pressure." Think of two old deserted barns of exactly the same size, which stand on either side of some old country road. Suppose both barns are filled with barn swallows, with swallows sweeping to the eaves across the hay, by the old stalls, and through the harness room. Each barn would be filled with ever-moving life whose sum would be the sum of all the swallows in the barn. Change from the two barns full of swallows to two jars full of gases. A liter of chlorine gas standing beside a liter of hydrogen gas is, in our belief, filled with moving life. The molecules which constitute the matter of the gas are in constant motion, bounding and vibrating to and fro within the confines of the vapor. There are so many in even a liter of gas that our finite minds can scarce imagine their number. Think, then, what a wonderful imagination a man must have had to conceive that every liter or other fixed volume of a gas contains exactly the same number of these tiny bodies. If, for example, we could count (as we cannot) ten billion molecules in a cubic centimeter of sulphur dioxide gas we should find just ten billion molecules in a cubic centimeter of carbon dioxide gas and the same amount in a cubic centimeter of nitrogen gas. One condition must always be fulfilled. The measured gases must always be at the same temperature and pressure. It is only under such circumstances that the hypothesis of Avogadro applies. That one condition accomplished, we believe that every body of gas, be it wide as the world or small as an air bubble in a glass of water, contains the same number of molecules as any body of gas equal in size to itself.

The foundation of fact on which rests this widely believed and valuable hypothesis is too long to give in such a work as this. Some of the reasons why we trust in its truth will be developed later, but much of the data on this subject which has been laboriously gathered through a hundred years must be omitted.

The gas laws as a foundation stone of chemical theory. — The relation of gases to each other, the action of gases under varying conditions of temperature and pressure, their weights and their characteristics, stand in many ways as the basic facts of chemical theory. The reason for this may be found in the extraordinary similarity of behavior of these bodies. Among the changing forms of matter this one great class stands apart and alone. Every member of this group obeys a whole series of definite laws. By that obedience they blaze out a path opening to constantly greater and fuller knowledge, not only of this single state of matter but of all the manifestations of the natural world with which they come in contact.

CHAPTER XVIII

THE SOIL

The four elements of the ancients. — To the ancients there were four elements, air, earth, fire, and water. To us not one of the four is a true element. We know that water is a compound of hydrogen and oxygen, that fire is generally produced by the combining of oxygen with some oxidizable substance, that air is a mixture of oxygen, nitrogen, and other gases. The ancients were far enough from the truth in their conception of these three of their four elements, but when they called the fourth, earth, a simple substance, they went even farther astray.

Natural resources hidden in the crust of the earth. — Every metal and every non-metal is found somewhere in this crust of our earth, and each bears some part in the economy of nature, while many of them serve a thousand diverse uses. Most of these have been brought out by the invention of man in the last hundred years or so, since comparatively few of the earth's natural resources were developed by man's ingenuity during the two or three thousand years that preceded this era. Hoards of natural wealth have been dragged into service and harnessed to civilization's advance, by the gradual increase in invention, by man's outreaching to overcome barriers which stood between him and the life he wished to live.

The soil which nourishes the crops. — The soil, however, this very topmost layer of the earth whereon we walk and ride, which carries our houses, our wagons, and our crops, has a use as old as any which has existed since this earth became a habitable sphere — a use as old as that of air and of water, namely, the production of food for man and beast.

The value of the soil. — No matter how far our life may lead us from farm and field, it is on the products of the soil that we must depend just as much as did primeval man, thousands of years ago. Bird and beast, reptile and insect, like man, depend directly or indirectly for their life on the plant world whose existence in turn depends on the food it can take from the soil.

Sowing

The ancient use of soil for the production of the food of man and beast.

Elements of the soil. — Barring carbon, which they obtain from the carbon dioxide of the air, oxygen, and possibly in a few cases water, the vegetable organisms of the soil generally can use nothing for their growth save what they take from the earth in which they stand. Most of them have no means of absorbing the uncombined nitrogen of the air, though many of them require large quantities of nitrogen which they take in combined form. The elements of the soil which are essential to the growth of the great body of plant life are sulphur, phosphorus, potassium, calcium, magnesium, and iron, sodium, silicon, and chlorine. Small traces of other elements occasionally occur and the hydrogen and oxygen, which plant life needs so abundantly, come from the water of the soil.

Fertile and sterile soils. — Since plants require such varied foods, it is evident that the soil must be made up of many things. Indeed, its fertility depends upon the presence of many different compounds of the elements mentioned above. If such compounds are present in large quantities the soil is fertile. If in small amounts it is sterile or barren. A fertile soil unmanured, and one from which crops are removed year after year, becomes gradually more and more sterile, until, its food elements exhausted, it will no longer support plant life.

American Pastures
The soil as a producer of food for plant and animal.

The use of manures. — The fertilization of the ground by manures is a chemical process centuries on centuries old. Manures may act in one or both of two ways, either supply plant food directly, or so change some compound present in the soil as to render it available where before it could not be used. The three elements which such fertilizers most need to supply are nitrogen, phosphorus, and potassium, for these are the three most generally used and most readily exhausted. The great piles of farm manure which stand in fields and barnyards in the fall contain large quantities of these three important elements. Such piles are left outside for months to decay in order that these substances may be made more available. The farmer knows that well-rotted manure is of more value as a fertilizer than fresh manure, but he does not always realize that the reason for this lies in the fact that the stacked fertilizer is losing its useless carbon as carbon dioxide and marsh gas, and is gradually changing its nitrogen and phosphorus to a condition where it can be most easily used.

Different types of fertilizers. — Some types of seaweed give fertilizing value not far from that of farmyard manure, while guano is heavy in phosphates. Nitrates, especially sodium nitrate, are largely used to give needed nitrogen to the ground, and are in such demand for that purpose as to be an important article of commerce. Lime which provides calcium acts as an alkali, neutralizing the acids in the soil and setting free foods which would otherwise be locked fast in unavailable compounds.

The food that different plants require. — In fallow fields where the exhausted soil is left for some seasons untilled to recover its strength, the changes which make for fertility are brought about by the slow agency of air and rain, of frost and snow, which break down and build up the various necessities of plant existence. Rotation of crops, changing from wheat to clover, from turnips to barley, often renovates the soil amazingly. Such renovation is probably due to the fact that the different crops used in rotation require different foods. Wheat and barley require phosphorus. Onions and strawberries require potassium. In the change from crop to crop, one year perhaps requiring phosphorus, the next potassium, time is given for the fields to recover sufficient strength to impart each element required in due proportion.

The scientific study of the soil. — The natural variations of the soil, the fact that a soil rich in the food which one crop demands may starve a crop requiring other nourishment, has made the work of the chemist who determines the constituents of the soil, who makes soil analyses, of especial value. By finding what plant foods are present in various soils the analyst can determine not only what crops should grow well with the food provided, but also what fertilizers will bring the ground to its highest state of efficiency. The scientific study of the soil and its relation to plant life is bringing forth every year more and more bounteous harvests and is driving away the fear of general starvation which oppressed some of those philosophers who mused a century ago on the problem of increasing population and the diminishing of the fertility of the fields.

The soil is not by any means wholly or largely made up of the plant foods we have been discussing, even though these foods are among its most important parts, considering the necessary relation they bear to the life of all animate creation. The chief relation which the main part of the soil bears to plant life is its action in supporting the plant, just such an action as is performed whenever we pot a geranium. We take up enough soil with the plant to cover the roots and hold the plant mechanically upright. Far more than ninety per cent of the soil acts as such a support for vegetable life, serves as a distributor and reservoir for the water which such life requires, and slowly, very slowly, itself breaks down into a state where it can serve as plant food.

How earth is formed. — Earth is broken down, separated, or disintegrated rock. If the rock of the soil be limestone

Rock
The rock of Gibraltar.

the soil will be heavy in calcium. If apatite (the natural phosphate of lime) occur in the region, phosphorus as well as calcium will be found in the soil. Potash is present wherever granite rocks split and break down into their parts. Outside of calcium carbonate, the limestone whose action we have already discussed, and humus, the general term under which we may include the decaying organic matter of the soil, — outside of these two, the chief constituents of a soil are sand and clay. These two compounds we shall take up separately.

Sand, silicon dioxide, silica = SiO_2. — Stand beside the ocean at ebb tide when the receding waves have bared long gleaming miles of wet and glistening beach. Watch the eddying swirl of sand storms on the deserts of the western plains. Dig below the surface in a thousand places north and south, east and west, and see how soon you find the upper surface giving away to strata of white or brown sand. "As many as the sands of the desert." "Like sands on the shore of

SAND
The sand of the shore.

eternity." Over and over again in literature we find some writer repeating this conception of the great stretches of sandy wastes. It is not hard indeed for us to realize that silicon, of which sand is the oxide, is found in greater quantity than any element of the mineral world.

Gems which contain silica. — White sand is ordinarily almost pure silicon dioxide or silica; quartz is silica, rock crystal being the purest form of silica known; while many rare crystals and precious stones such as the amethyst and jasper, the onyx and the opal, are composed chiefly of this compound.

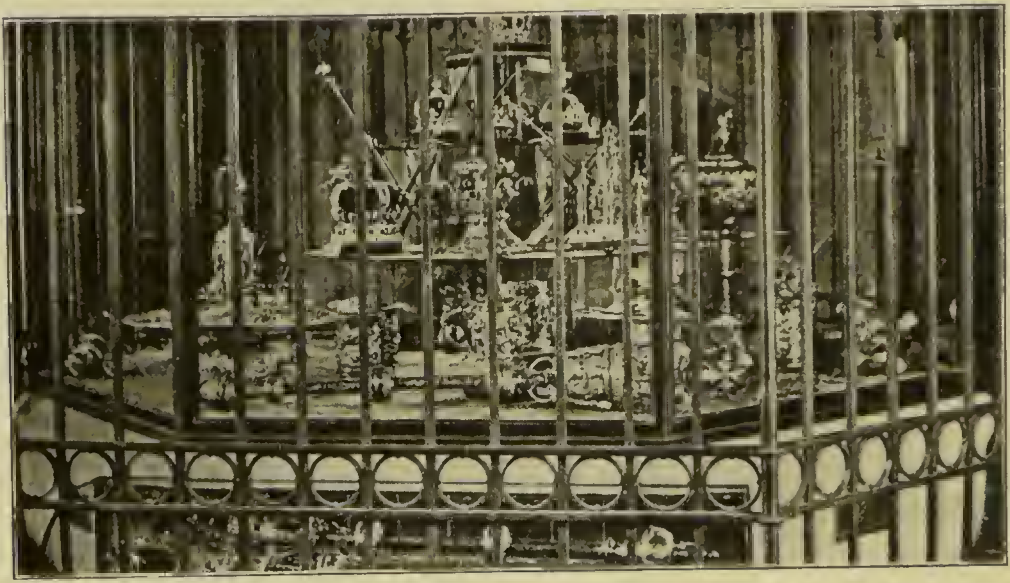
Flint. — "Hard as flint," says the proverb, and it might read, "Hard as silica," for flint is a type of silica, and its hardness is so great that, when struck by steel, it gives off red-hot particles of metal which set fire to tinder or to gunpowder. This was the action by which the old flint and steel of the flintlock gun set fire to its charge.

Properties of silica. — Since silica serves as the main constituent of the soil it is evident that it needs to be practically insoluble. Though insoluble in pure water it is not insoluble in some solutions. It will dissolve to a certain degree in alkalies, and the result of its solution is found in the hard outer coating of various grasses and cereals. The shining armor of the bamboo and the stalk that bears the nodding blade of wheat both hold their shape largely because of the silica in their composition. Sword grass gets its cutting edge from the same compound, while even the stem of the potato contains a considerable percentage of this compound of the soil.

Clay, aluminum oxide. — When we come to consider the last great constituent of the soil, its clay, we find that we do not lose the silica, we only add a new element to it. Clay contains from fifty to sixty-five per cent of silica or oxide of silica, and from twenty to thirty-five per cent of alumina or oxide of aluminum (Al_2O_3). It generally contains some water, some oxide of iron, a little magnesia, and sometimes potash or soda, but it is chiefly a silicate of alumina, and the three elements which give it its characteristic properties are aluminum, silicon, and oxygen. The aluminum differs from most of the other elements of the soil in one respect. It is almost never found in the plants which spring from the earth. It remains steadfast to its place in the substructure which bears the vegetable world. Only in a few of those low types of plant life which are classed as cryptogams is it at all a common discovery.

Quartz

Granite.—The granite of the unyielding hills, such as underlie the rocky farms of New England, is one of the oldest of the substances which form the crust of the earth. This rock is no simple substance, but a mixture made up in varying proportions of quartz, feldspar, and mica, and its different colors come from small traces of iron and manganese compounds. The white part of granite, the white streak, is the quartz, and this is silica, the same silica which we have been discussing. Feldspar, the cream-colored opaque part, is a combination of silica of aluminum and potassium, while mica, the glittering scales that peel from the rocky stone, also contains silica and alumina and potash (the carbonate of potassium). The alumina and potash may be replaced by oxides of iron and magnesium. Year after year, and century after century, solid rock is beaten by snow, rain, and hail and split by the glittering wedges of ice when the crannied water freezes and expands with terrific force. Little by little such phenomena of summer heat and winter cold break a rock down into its component parts, while rain-water, containing carbonic acid, dissolves away those parts of the structure (such as the silicates of alumina) which are soluble in this liquid. These are washed away from the denser quartz and form beds of clay which are chiefly composed of silicate of aluminum. Kaolin, the potter's clay, is the purest form of clay that we have, containing the most alumina (oxide of aluminum). Other types found in the earth which are likewise silicates of aluminum are fullers' earth and pumice-stone. Loam contains much clay and some of the paints that are dug from the ground, ocher, umber and sienna, are clays colored by various metals. The clay or slate of the ground gives alum, that important substance used widely in commerce.



JEWELS

Jewels of the British Crown.

Jewels and tilled fields. — The hidden gems of the soil, the precious stones, ruby, sapphire, opal, and amethyst, are simply different types of these substances which make up the earth's crust. The man searching the deep mine or the broken field for these rare things of beauty finds substances which have been converted by the crucible of molten earth into gems. These hard crystals, these transparent jewels, have but a hundred thousandth part of the value that the good brown earth has for us, but a millionth part the place in the great scheme of nature that the soil holds — that mother of us all, that spreading field from which comes the truest wealth of all the nations, the produce of the land.

CHAPTER XIX

SULPHUR

Symbol for Sulphur = S

Occurrence of sulphur. — To many of the ancient people who lived in southern lands, the interior of the earth was a place of unknown horrors. They firmly believed that beneath the thin crust on which they lived were infernal regions filled with "fire and brimstone," and teeming with fearful mysteries. That such was the conception of the people in southern Italy and Sicily is not strange since they possessed in their very midst volcanoes like Vesuvius, and knew, by constant experience, of earthquakes, of slight tremors of the earth, and of the rising of sulphurous fumes. Sicily especially has always held great wealth of natural sulphur, and from early times to now a very great portion of the world's supply of the yellow element has come from the volcanic region which lies down near the toe of the boot that forms Italy. Ship after ship came sailing into our harbors laden with the yellow sulphur of Sicily, some of it in the form of the roll sulphur or brimstone, which appears in shiny sticks, and some of it in the shape of powdered sulphur, or flowers of sulphur, two types which differ only because of the different methods of their manufacture. To-day Louisiana has outstripped Sicilian production because of the discovery of modern methods of working.

How sulphur is found in the earth's crust. — For years the trade in Sicilian sulphur was enormous and profitable, both because of the great plenty of the element thoughout that region and because of the cheapness of labor there, but other deposits which have more recently come to light are being worked to good advantage. Japan is sending many cargoes to the Occident, and a deposit in one of the islands of the New Hebrides is sending yellow loads from the southern Pacific to northern lands. In North America sulphur is found largely in the West, in Texas, California, and the Yellowstone, where sulphur springs abound. Mexico has also developed a considerable export trade. The coal and sulphur, iron, copper, and the like, which are hidden below the upper soil, appear in various ways, and in many different kinds of deposits. Many of them appear in strata, arranged like layer cake or Washington pie. Think of a cake where two thicknesses of chocolate cake are outside and one of yellow cake between. That would fairly represent a stratum of sulphur between two strata of clay. In the earth's surface we often find sulphur occurring

A Natural Crucible

A volcano sending forth steam and sulphurous fumes.

in layers placed between strata of different composition and forming a whole which is called a stratified formation. Not all the sulphur is found stratified with rock or clay, but much of it occurs in the so-called "living" beds, deposits in which the sulphur earths are constantly decomposed in those great boiling kettles of nature which occur in volcanic lands. These living beds, to which the great forces of nature incessantly furnish heat, are types of one of the most impressive of nature's processes, the work of the volcanoes and their allies the hot springs, fumaroles and geysers, all of which give outward sign of inward disturbances. More than one substance which occurs nowhere else in nature is found in the compounds and elements sent forth from those natural laboratories whose reactions are mightier than the hand of man ever has or apparently ever can produce.

Where sulphur appears in the organic world. — From volcanoes to egg or mustard spoons seems a long step, yet the blackened surface of a spoon which has been used for eating eggs is due to the sulphur in the yolk of the egg, and for the blackened mustard-spoon the sulphur of the yellow mustard is responsible. The tarnishing of silver or copper comes largely from the sulphur commonly present in the air of the house. Most coal contains some sulphur, and, generally speaking, the poorer the coal the more the sulphur, so that the leaking gases from the kitchen stove or the furnace send out sufficient of the element to tarnish any of the metals which readily combine with sulphur that may be near. Coal gas, the ordinary illuminating gas, adds its quota of sulphur in combination, moreover, to the air of any room where a gas light burns. White lead, the commonest of all white paints, is blackened by sulphur and lead sulphide is formed.

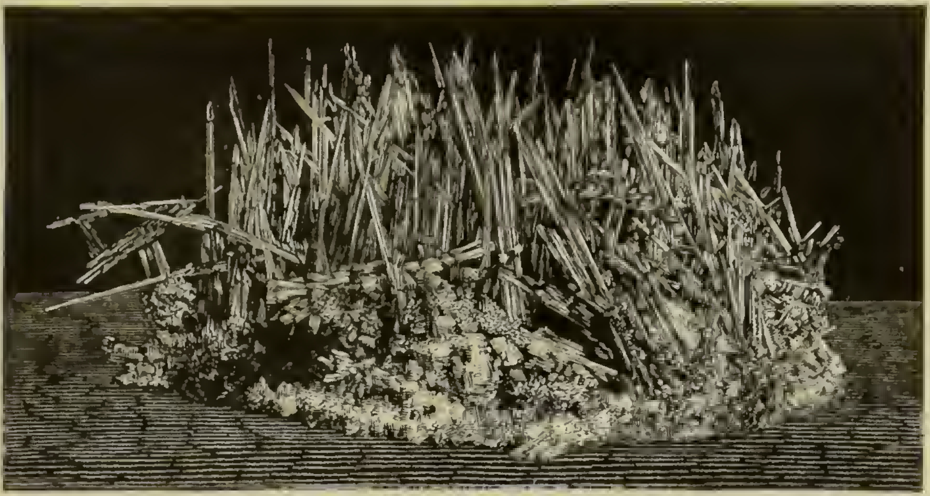
Eggs and mustard are not the only foods which contain sulphur; several of the vegetables, cabbage and beans especially, contain varying amounts. Since sulphur is a constituent of such common foods it is by no means surprising to find that sulphur makes up quite a proportion of the solids of the human body, more than four ounces of it being present in the makeup of a man of ordinary size.

The sulphides. — Of all the elements sulphur is most like oxygen, so it is not strange to find that it is, like its near relative, a great combiner. Indeed, nothing characterizes it more than this one property of combination. Metal after metal is found in the form of sulphides, that is of compounds of sulphur, and these sulphides are found all over the world. In rocky heights, near city streets, and northern mountains, lead sulphide or galena is found. On the shores of Lake Superior copper sulphide is found, as copper glance or copper pyrites (Cu_2S), while mines of old Spain and of California (the new Spain of which the grandees dreamt) contain cinnabar, the sulphide of mercury (HgS). One sulphide found in the greatest plenty and scattered over a great part of the world is iron sulphide, better known as iron pyrites (FeS_2) or "fools' gold." As we shall shortly see, this last sulphide has immense importance as the source of the commercial necessity, sulphuric acid, the chief acid of sulphur.

Properties of sulphur. — The properties of sulphur have much to do with its preparation and use. It is one of the solid elements; at ordinary temperatures is pale yellow; is insoluble in water, but readily dissolves in a sulphide of carbon, carbon bisulphide; does not conduct electricity and scarcely conducts heat. It is a light solid, being only about twice as heavy as water. So much for its general characteristics. Now for its special ones. First and foremost comes the ease with which it melts, turns into a gas and burns, combining with oxygen. The ease with which sulphur changes its state with heat is the cause for its wide-extended use in matches, in gunpowder, in fireworks. Sulphur indeed melts at a temperature just above that of boiling water, but before it becomes a gas it passes through a series of changes quite different from those experienced by any other element.

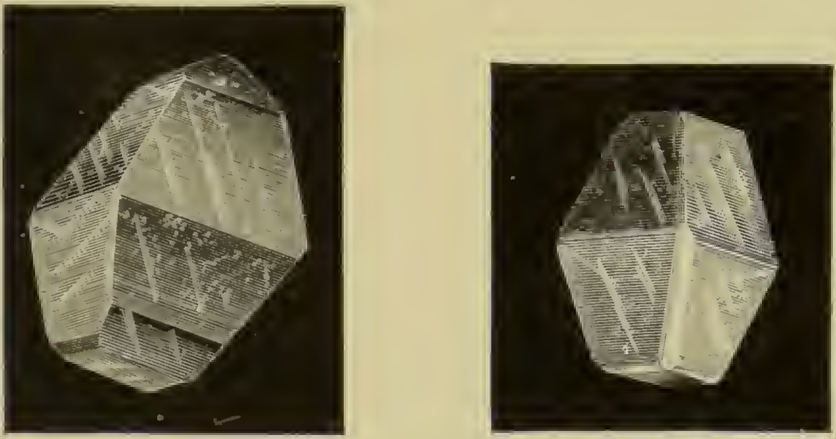
Sulphur in the three states of matter. — When sulphur passes 114.5 degrees C., that is, when its temperature is slightly above the boiling-point of water, it turns to a clear amber-colored liquid which flows easily and can be readily poured from one vessel to another. As the temperature rises above that point the clear amber deepens to a dark red and the liquid thickens, until, at 220 degrees C., the color is almost black and the mass is so thick and sluggish that the vessel in which the heated sulphur lies can be turned upside down without its running out. As the temperature rises still higher, the color remains the same, but the substance now becomes liquid once more, though not the easily flowing liquid of the lower temperature. At 448 degrees C. the liquid boils and a pale yellow brown gas comes off which condenses back to the yellow solid sulphur with which we started, while, at the same time, if the sulphur is melting in the air, a part of it takes fire and, burning with a blue flame, forms sulphur dioxide. Here is a whole cycle of physical changes all of them dependent on heat and none of them involving the chemical structure of the element. The solid sulphur of the start, the clear amber fluid of the first liquid, the sluggish viscous mass which came with greater heating, the darker liquid as well as the light yellow brown gas, all were sulphur and nothing else. The only chemical change occurred when sulphur, as it burned with a blue flame, turned to sulphur dioxide ($S + O_2 = SO_2$).

The allotropes of sulphur. — Sulphur has more than one face, appears under more than one form. As oxygen has an alias in ozone, so sulphur appears likewise in transformation, but where oxygen has only one alias in ozone, sulphur appears in four different forms or allotropic modifications, each of which is sulphur and nothing but sulphur, though each has a distinctly different physical appearance. These four are prismatic sulphur, rhombic sulphur, plastic sulphur, and white amorphous sulphur. If you think of a thicket of spears such as the men at arms carried at Agincourt, or Crecy, in the wars when England was trying to gain possession of France, you will have no bad idea of the way that prismatic sulphur, the first of the different kinds of sulphur, looks when it is first prepared. Long translucent needle-like spears are these prisms of prismatic sulphur. They are prepared by heating sulphur in a beaker till it is thoroughly melted, and then allowing the mass to cool until a part and a part only is solidified. When the liquid portion is poured off from such a mixture, the solid part is found in this spearlike form lining the vessel.



Prismatic Sulphur

Rhombic sulphur. — Precious stones cut in smooth-faceted planes sometimes show forms similar to that of the other and more permanent form of crystallized sulphur, orthorhombic or "rhombic" sulphur. This is the form to which prismatic



Rhombic Sulphur

sulphur tends to change, as this first type mentioned endures but a short time at ordinary temperatures and is never found in nature. Native sulphur, on the other hand, is commonly found in the little rhombic crystals which have such smooth facets that they might well represent smooth building blocks.

Plastic sulphur. — When chewing gum has been well chewed and reaches the point where it is thoroughly elastic, or when rubber is in form of rough black elastic masses, we have substances which resemble outwardly the third form of sulphur, plastic sulphur. If melted sulphur near the boiling-point is poured into a dish of cold water, this sulphur, wholly unlike any sulphur you have ever seen, drops into the dish in a coiling elastic mass of amber tint that you can compress and pull out almost as if it were India rubber. This allotrope is the lightest of all the forms of the element and is most soluble in the common solvent for sulphur, carbon bisulphide. It soon turns into the first and commonest form, the rhombic, which we have just considered.

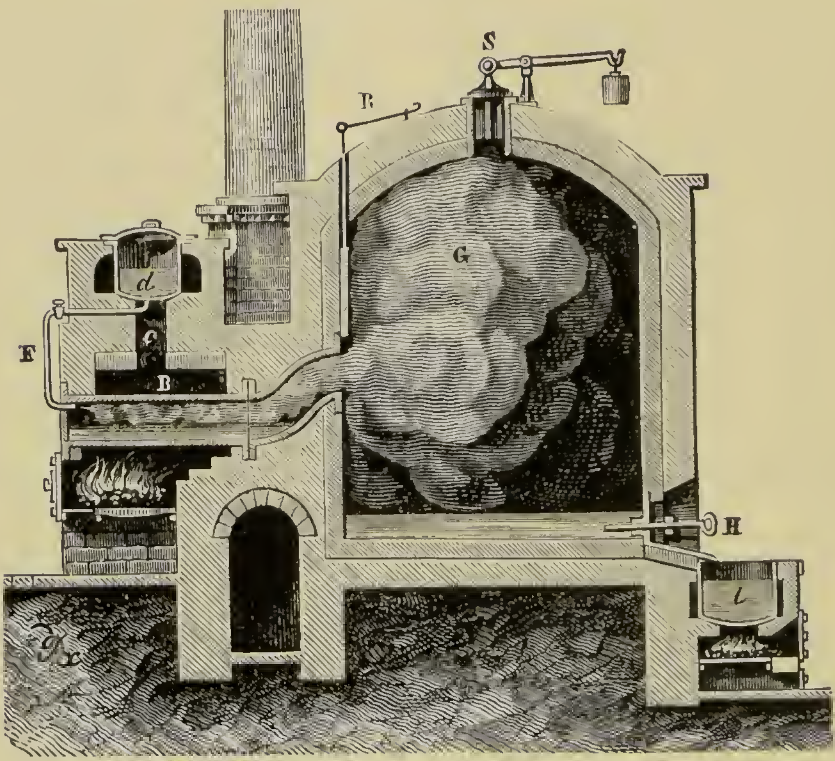
Flowers of sulphur. — On the drug store or the chemist's shelf may be seen the commonest form of sulphur that we know, the flowers of sulphur. The form is rhombic, but when these flowers of sulphur are produced there is formed with them a small amount of what is known as white amorphous sulphur, almost milk-white in color and sometimes appearing when sulphides in solution are acted on by some type of an oxidizing agent. When an oxidizing agent is present, a white insoluble precipitate may fall which has given many a beginner trouble due to his thinking that he had a white metallic precipitate when he really had white amorphous sulphur, a substance without crystalline shape, differing in that respect from the two earlier crystalline allotropes.

Milk of sulphur. — Milk of sulphur does not occur in nature, nor does it belong among the regularly recognized allotropes of sulphur. It is produced as a precipitate when certain sulphides, heavily charged with sulphur and called polysulphides, are treated with hydrochloric acid. When such action occurs, a fine white powder (milk of sulphur) falls.

Tests for sulphur. — The test for sulphur, speaking roughly, is its appearance and its action, both wholly unlike any other element or compound, or speaking more exactly, the fact that when it burns it forms sulphur dioxide, burning with a pale blue flame. That oxide added to water gives sulphurous acid, and the tests for the acid formed will be discussed in the succeeding chapters. The combination of sulphur with metals to form sulphides is another test. In many cases such a union, like the union of metals with pure oxygen, is a true combustion, giving off light and heat. So heated copper wire thrust into sulphur vapor or in contact with solid sulphur burns brilliantly, forming copper sulphide ($Cu + S = CuS$); hot iron does the same, forming an iron sulphide ($Fe + S_2 = FeS_2$).

Purification of sulphur. — Sulphur is practically always mixed with earthy substances, various kinds of dirt and soil. These impurities have to be removed from the element before it can be available for use. In Sicily, where, as we have said, sulphur is most plentiful and labor is cheap, workmen dig out earth and sulphur together and simply cover the mixture with earth and sods. Then they set fire to the whole mass. Most of the sulphur melts and runs out into troughs. No small quantity of the sulphur is, however, burned by this process, when the resulting sulphur dioxide is lost completely. In the more modern processes this loss of sulphur is prevented by heating the substance out of contact with the air. Every process of manufacturing sulphur whether old or new depends first on the heating of the sulphur, second on its separation from the other substances around, because of the lower melting-point of the element, or third because of its conversion into a gas which is afterwards condensed.

Manufacturing uses of sulphur. — When most elastic rubber bands are placed around gold or silver ware, black lines show where the bands were placed. Under these circumstances gold or silver sulphide is formed and the sulphur of the rubber band unites with the metal. A large part of the refined sulphur manufactured goes into vulcanizing rubber. When the cartridges begin to crackle in the hunting season of October, sulphur is contained in much of the gunpowder that sends the bullet forth. The sulphur on the so-called sulphur match shows probably the best known every-day use of this element. The yellow sulphur of the common match has the mission of carrying on the flame from the phosphorus of the tip to the wood of the match splint. The tip of the match often contains phosphorus, an element that ignites very easily with friction when the match is struck. The heat of that ignition sets fire in turn to the sulphur which bubbles up, melts, and



Purification of Sulphur

One of the older forms of apparatus for purifying sulphur by distillation is shown here. The sulphur, melted in *d*, passes through F to the retort B. There it boils and the vapor enters the chamber G. Here it condenses to the solid flowers of sulphur.

burns, turning to sulphur dioxide. The sulphur as it burns sets fire to the wood.

Sulphur as a disinfectant. — The uses of sulphur in its acids will be left to another chapter, but its services as a disinfectant should be mentioned here. Germ life dies in the fumes of sulphur dioxide. In consequence the burning of sulphur has long been practised to cleanse rooms from dangerous germs which might spread contagion, or to fumigate places where infectious disease could lurk. While some of the organic substances are taking the place of sulphur as disinfectants in sick rooms to-day, the element still remains of great use in clearing vineyards from an attacking fungus which the grape raiser needs constantly to combat. As a result a large portion of the world's sulphur goes to protect the grapes of the vineyards which lie on the sunny slopes of the continent of Europe.

The general uses of sulphur. — Part of our daily foods, builder of our body cells, essential part of two great acids and their compounds, incorporated in the rubber that serves a thousand purposes, destroyer of germ life, combiner with metals holding the wealth of the earth in store, all in all, sulphur ranks with the greatest of non-metals. Few elements do more to serve mankind.

CHAPTER XX

SULPHUR DIOXIDE (SO_2) — SULPHUR TRIOXIDE (SO_3) HYDROGEN SULPHIDE (H_2S)

Sulphur dioxide, the mild bleacher. — When your last summer's straw hat went into the closet for its winter rest, the white glistening straw of the spring had turned to a dull yellow. Contrary to general opinion the summer's sun and shower did not bring that color. It was there at first. The white straw had been bleached, and the bleacher which gave it its fresh appearance was sulphur dioxide gas. Chlorine is the energetic whitener. Sulphur dioxide is the mild bleacher which whitens flannels and straw, basket-work and hair, not by eating out the color of their structure but by combining with the coloring matters so as to form colorless compounds. Water and light, light especially, break down these compounds in time, and straw hats grow yellow, flannels change their hue, and baskets grow dark and old as the colorless compounds break up and the color which was there at first disappears. Sulphur dioxide is a valuable bleach because of the comparatively slight harm which it does to fabrics, but it is a comparatively temporary bleach because it only hides and does not remove the color.

Occurrence of sulphur dioxide. — When sulphur burns in oxygen sulphur dioxide is always produced. The sulphur smell of coal gas and the odor of a brimstone match both come from the formation of this compound. Its traces may be found in a thousand places. Quick as sulphur is to unite with oxygen, once it has combined the dioxide of the element will not allow any other body to combine with oxygen in its presence. Thrust a match into a bottle full of sulphur dioxide and it is instantly extinguished. Wherever sulphurous vapors escape, the life of vegetation which demands oxygen dies. In uncleaned chimneys soot gathering year by year sometimes takes fire. Few better means of extinguishing that fire can be found than burning sulphur on a shovel in the fire below, where the draft will take the gas formed up the chimney. When this is done the fire soon stops. The rising sulphur dioxide puts out the flame. It will allow no oxidation in its presence.

Sulphur dioxide as a germ destroyer. — The last fact explains the use of burning sulphur as a disinfectant and a destroyer of germ life. Many of these infinitesimal organisms, known commonly as germs or microbes, require oxygen for their existence as much as we do, and die suffocated when a burning sulphur candle throws off clouds of sulphur dioxide gas into the air. Valuable as this principle is in many cases, one use, its employment for food preservation, is at least questionable. Yet since considerable quantities of sulphur dioxide are required for this purpose, mention of this commercial usage of the gas should be recorded here. On this principle of germ destruction rests the disinfectant or germicidal employment of all compounds of sulphur.

Common uses of sulphur dioxide. — Some other uses of this important oxide gas, such as its occasional employment in the making of artificial ice and the preparation of pulp for paper, might be mentioned, were it not for the fact that the first of these processes has already been considered and the second is to be discussed later on. In consequence we can turn directly to the preparation and properties of this common oxide of sulphur.

Properties of sulphur dioxide. — We have already seen that burning sulphur in the air produces sulphur dioxide. That is an easy and simple method of preparation. If we burn sulphur in a bottle full of air, and, before the heavy sulphur dioxide gas can escape, turn the receiver quickly over and place it mouth down on a basin of water, we shall see a slow change.

Little by little the water begins to creep upward into the bottle, rising higher and higher. It is taking the place of the sulphur dioxide in the bottle. As the gas goes into solution the water rises until all the gas has been dissolved. Evidently one of the chief characteristics of this compound is its solubility. We commonly confuse the smell of sulphur dioxide gas with the smell of sulphur, for the only time we get any strong odor from sulphur is during its burning, when it changes to this gaseous oxide. The odor of the sulphur match gives the characteristic smell of sulphur dioxide. The gas is far heavier than air, is colorless and readily condenses to a clear liquid at the zero-point of our Fahrenheit thermometer scale. Its power of turning with comparative ease from a gas to a liquid and vice versa is the one characteristic which makes it available for use in manufacturing artificial ice. This is a similar action to that of ammonia already noticed.

Preparation of sulphur dioxide in the laboratory. — Making sulphur dioxide from burning sulphur is not wholly practicable as a laboratory process, as the burning sulphur will not give a steady stream of gas. In consequence for this preparation we take advantage of the fact that sulphur dioxide is evolved when heated sulphuric acid acts on metallic copper. In this reaction two parts of hydrogen sulphate (sulphuric acid) added to copper produces copper sulphate, water, and sulphur dioxide.

Hydrogen sulphate + copper = copper sulphate + sulphur dioxide + water.

$$Cu + 2H_2SO_4 = CuSO_4 + SO_2 + 2H_2O.$$

One part of sulphate makes copper sulphate, the other makes sulphur dioxide and gives up its excess oxygen to form water.

Bleaching with sulphur dioxide. — From first to last sulphur dioxide is a bleacher. As the snow comes with its white blanket covering gray stone, red clay, and brown earth, turning them all to dazzling white, so sulphur dioxide let loose upon a garden would bleach petal of red rose and of blue violet to the same dull white. All flowers fade in its presence, though they retain their original form and shape.

Sulphur trioxide the theoretical basis of sulphuric acid. — Of rare occurrence outside the laboratory, of but little commercial interest in itself, the second oxide of sulphur, sulphur trioxide, is in one respect far more important scientifically than the oxide that we have just considered. This is the compound from which comes the most widely used of the great group of acids, sulphuric acid, a compound used for one purpose or another, wherever the tall chimneys of manufactories rise, and men, gathering in communities, acquire material wealth. Half of the next chapter will be given over to this acid. Our discussion here is confined to the oxide whence it comes.

The acid-forming properties of the non-metallic gaseous oxides. Anhydrides. — One of the most striking characteristics which the gaseous oxides of the non-metals share in common is their power to form acids when mixed with water. The oxides of carbon, nitrogen, sulphur, and phosphorus all share this power, and when united with water give as results carbonic, nitric, and nitrous, sulphurous and sulphuric, phosphorous and phosphoric, acids. In all these cases the mere addition of the oxides to water produces a new compound which contains the hydrogen characteristic of the acid. Such non-metallic acid-forming oxides are called acid anhydrides; acid because of their acid-forming powers, and anhydride, meaning without water, because it is only when the acid-forming hydrogen present in the water is removed that we get the non-metallic oxide, the acid anhydride. Now sulphurous anhydride (sulphur dioxide) (SO_2) is a compound of sulphur which contains less matter, one less atom of oxygen, than sulphur trioxide (sulphuric anhydride) (SO_3). In general we use that ending ous to indicate a lower form or one which contains less water. None of these acids show more clearly their formation than these two acids of sulphur, sulphurous which comes from sulphur dioxide and sulphuric which comes from sulphur trioxide.

Sulphurous acid made up of hydrogen and a radical called the sulphite radical. — If we run sulphur dioxide gas into water, hydrogen oxide, we get a new compound, an acid wholly distinct from the two oxides with which we started. Sulphur dioxide contains one atom of sulphur and two of oxygen. Hydrogen oxide contains two atoms of hydrogen and one of oxygen. Add them together and you have this result:

2 hydrogen atoms		1 oxygen atom
	1 sulphur atom	2 oxygen atoms
2 hydrogen atoms	1 sulphur atom	3 oxygen atoms

$$SO_2 + H_2O = H_2SO_3.$$

This acid which is formed from sulphurous anhydride we call sulphurous acid, because it contains one less atom of oxygen than sulphuric acid formed from sulphuric anhydride. If we examine this acid carefully we shall find that it acts in a striking way. Under most circumstances it does not tend to go back to the sulphur dioxide and water which formed it. It does not decompose into its three elements, hydrogen, sulphur, and oxygen. Instead we find it breaking sharply into two parts, one made up of the two hydrogen atoms (which we may call the hydrogen part) and the other made up of the sulphur atom and the three oxygen atoms. The second part made up of sulphur and oxygen sticks together under almost all circumstances. Bring the acid up to a metal and the hydrogen may fly off, but the other part hangs together. Under almost every condition this little group which we have formed acts in unison. Such a group as this (there are a number of them) we call a radical. Defining it more exactly we say that a radical is a group of atoms acting as a single atom. In this special case we may say that we have here the radical of sulphurous acid (SO_3) called the sulphite radical, a name derived from hydrogen sulphite, which is a second name for sulphurous acid, given it to distinguish it yet more sharply from sulphuric acid, hydrogen sulphate. The ending –ite belongs with the ending –ous, and means lower, while the ending –ate, belongs with the ending –ic, meaning higher. This radical, then, we call the sulphite radical, and wherever it appears in combination with metals we call those compounds sulphites.

Sulphuric acid made up of hydrogen and a radical called the sulphate radical. — Sulphur trioxide follows sulphur dioxide in its action. Look at this addition:

2 hydrogen atoms		1 oxygen atom
	1 sulphur atom	3 oxygen atoms
2 hydrogen atoms	1 sulphur atom	4 oxygen atoms

$$H_2O + SO_3 = H_2SO_4.$$

When we add sulphuric anhydride, with its one atom of sulphur and three atoms of oxygen, to water, with its two hydrogen atoms and one oxygen atom, we get an acid made up of two parts — one the hydrogen part and the other a radical part made up of a sulphur atom and four oxygen atoms called the sulphate radical. When sulphuric acid takes part in most chemical reactions this group of atoms stays together and works together. The acid of which it is a part is called hydrogen sulphate as well as sulphuric acid, and the compounds which the sulphate radical forms with metals are called sulphates.

The preparation and properties of sulphur trioxide. — To turn sulphur dioxide into sulphur trioxide a catalytic process is used similar in action to that used in preparing oxygen. If we pass sulphur dioxide over hot finely divided platinum used as a catalytic agent the platinum enables the dioxide to take up a third oxygen atom from the air and become sulphur trioxide. This forms sulphuric acid when left uncovered, because of its great tendency to unite with the water vapor of the atmosphere. The thirst for water which sulphur trioxide shows has a special name which we have met before, hygroscopic, which in itself means "eager for water." From the attraction which various substances have for water and from their power of absorbing or combining with it, many valuable chemical processes are derived.

Occurrence of hydrogen sulphide gas. — Next to the oxides of sulphur comes the gaseous compound of hydrogen and sulphur known variously as sulphuretted hydrogen, hydrogen sulphide, and hydrosulphuric acid. When an egg becomes bad the sulphur in its yolk combines with a part of the hydrogen present and sulphuretted hydrogen is formed. The smell of an old egg gives the characteristic odor of the gas. This odor may often be perceived rising from poor coal, especially in a poor draft, and sometimes may be perceived in the gas flame. This gas rises like sulphur from the bubbling pot of volcanoes, and is found, again like sulphur, in hot springs in volcanic lands.

Formation of sulphides. — Most important of all the properties of this compound is its power to form sulphides of the metals. When a silver spoon is tarnished by the sulphur in the air, hydrogen sulphide is at the bottom of much of the tarnishing, for this compound forms silver sulphide (Ag_2S) as it does the sulphides of lead (PbS), copper (CuS), tin, and several other metals.

The separation of the metals by the use of hydrogen sulphide. — The formation of these sulphides gives us a simple and effective means of separating the metallic elements which form them from solutions in which they are mixed with other metals. For example, suppose we have lead, iron, and sodium, in a solution together. We wish to remove the lead. If we pass a stream of hydrogen sulphide gas through the acid solution, the sulphide of lead, a black solid will precipitate and fall to the bottom, leaving behind the two other metals. Because of such action hydrogen sulphide has become a valuable separator of metals in combination. The reaction which occurs when hydrogen sulphide is added to any metallic compound would be similar to that which takes place when it is added to copper nitrate. This may be written as follows:

Copper nitrate + hydrogen sulphide = copper sulphide + hydrogen nitrate.

$$Cu(NO_3)_2 + H_2S = CuS + 2HNO_3.$$

The copper sulphide appears as a black solid while the nitric acid (hydrogen nitrate) remains in solution. To make sulphuretted hydrogen we can combine iron sulphide with sulphuric or hydrochloric acid. When this is done the sulphate radical of the acid unites with the iron while the sulphur of the iron sulphide unites with the hydrogen, forming hydrogen sulphide.

Iron sulphide + hydrogen sulphate = hydrogen sulphide gas + iron sulphate.

Hydrochloric acid is often used and the equations read as follows:

$$FeS + H_2SO_4 = FeSO_4 + H_2S.$$
$$FeS + 2HCl = FeCl_2 + H_2S.$$

When the diluted acid is poured on the sulphide (iron sulphide) we get a very rapid flow of gas. This is the common way of preparing the compound. But special precautions have to be taken to prevent its escape into the room. Sulphuretted hydrogen is slightly poisonous in quantities. It has a distinctly sickening odor and should only be used in a hood or draft closet, where a draft of air may carry it away. It is never wise to lean directly over a hydrogen sulphide generator and inhale the fumes.

Properties of hydrogen sulphide gas. — Hydrogen sulphide gas is colorless, moderately soluble in warm water, less soluble in cold, and heavier than air. Its solution in water turns blue litmus red. It burns with a blue flame, oxidizing its hydrogen to hydrogen oxide (water) and its sulphur to sulphur dioxide. Reaction for burning of hydrogen sulphide gas:

$$2H_2S + 3O_2 = 2SO_2 + 2H_2O.$$

Tests for hydrogen sulphide. — We have already said that white-painted houses as they age turn dark and black. That change is due in a large part to the fact that they are covered with a paint made up of white lead. White lead is a compound of lead known as basic lead carbonate, and lead sulphide forms readily from this compound. We shall consider this compound carefully in our study of lead. Lead sulphide is black where lead carbonate is white, so whenever sulphuretted hydrogen escapes in white-painted houses, or the fumes of sulphur come down from the chimneys, the white sides blacken. This tendency of lead compounds to change to black lead sulphide is used in a common test for hydrogen sulphide. A piece of paper is dipped in a lead acetate solution, and exposed to the stream of gas. If it turns black the gas is hydrogen sulphide. Lead sulphide has been formed.

CHAPTER XXI

SULPHUROUS ACID AND THE SULPHITES

Formula for sulphurous acid (H_2SO_3)

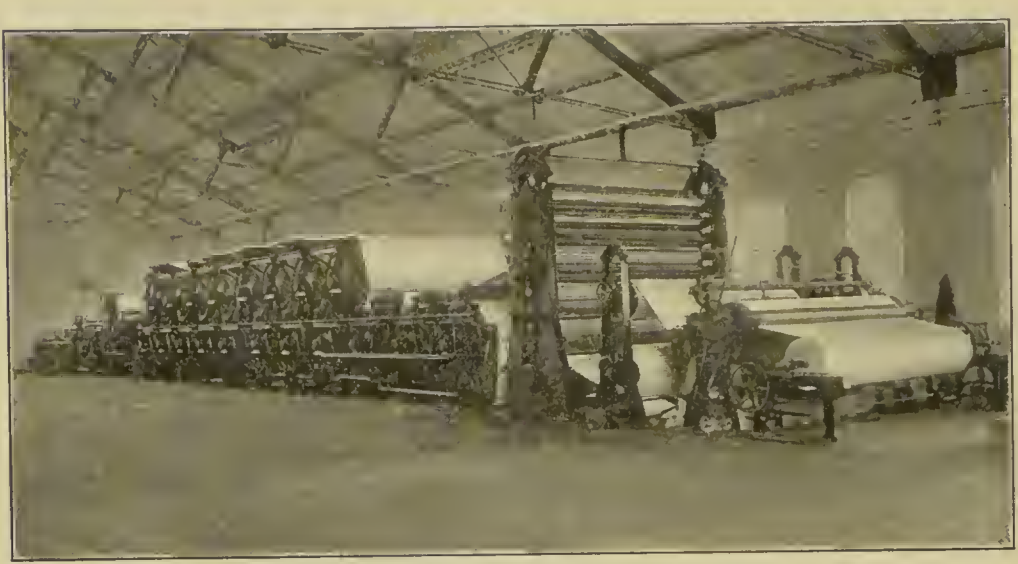
SULPHURIC ACID AND THE SULPHATES

Formula for sulphuric acid (H_2SO_4)

Sulphite pulp in paper. — When you see huge rolls of paper carted through the streets or unloading at the newspaper offices, when you pass high piles of wrapping paper at the store, do you wonder whence that paper came and how it was made? Most of the paper whereon we read the news of the day, most of the paper coverings to our parcels, comes from the forests of the north, from mountain-side and valley where the smooth poplar, the rustling hemlock and pine, the gleaming birch and the straight maple, stand. The greatest part of all the paper that is made in the world is made from wood pulp, and most of the pulp comes from the special kinds of trees we have mentioned. The chief thing that makes good paper possible at a low price is one of the acids of sulphur, sulphurous acid, which is chiefly of value because of its compounds, the sulphites. In this, as in so many other cases, chemistry makes possible a cheap and effective process.

Rags and wood pulp in paper-making. — For many years rags were used exclusively for paper-making, and even to-day many of the fine grades of paper are made from this material; but rags are comparatively costly, for linen and cotton must be prepared with care and passed through many processes before they can be manufactured into cloth and turned, after their usefulness in that form is done, to paper. Wood is far cheaper and can be far more readily turned to the requirements of paper manufacture. And the demand for paper exceeds all belief. Thousands and hundreds of thousands of tons are used the world over; to supply the need, forest after forest is cut down. In fact the demand for wood to make paper has grown so great that it is the chief cause of the devastation of our forests, a devastation which, if unchecked, will cause grave trouble before many years are passed because with the cutting off of the wood comes a consequent diminishing of the water supply, whose continuance depends largely on the forests which form great natural reservoirs, not only by the aid of the trees themselves, but also by the help of the earth mold formed by the decay of branch and leaf.

Other substances used in paper-making. — Besides wood and rags, other materials, straw, hemp, flax, jute, and esparto or Spanish grass, are used for this same purpose, but each in a comparatively small amount. We shall consider here only the commonest method of paper manufacture, its production from wood. Every substance used for paper-making must



A PAPER MACHINE. END VIEW

The continuous sheet of paper, passing over the long felt at the left, grows gradually stronger and stronger. It receives its final polishing in the calendars, the tall stack of rolls at the right and finally comes off in the huge paper rolls shown at the extreme right.

have, however, one property in common, the power of giving fibers, for paper consists of woody fiber matted into a sheet, and usually loaded with some earthy matter to give it weight, make it smooth and render it more waterproof.

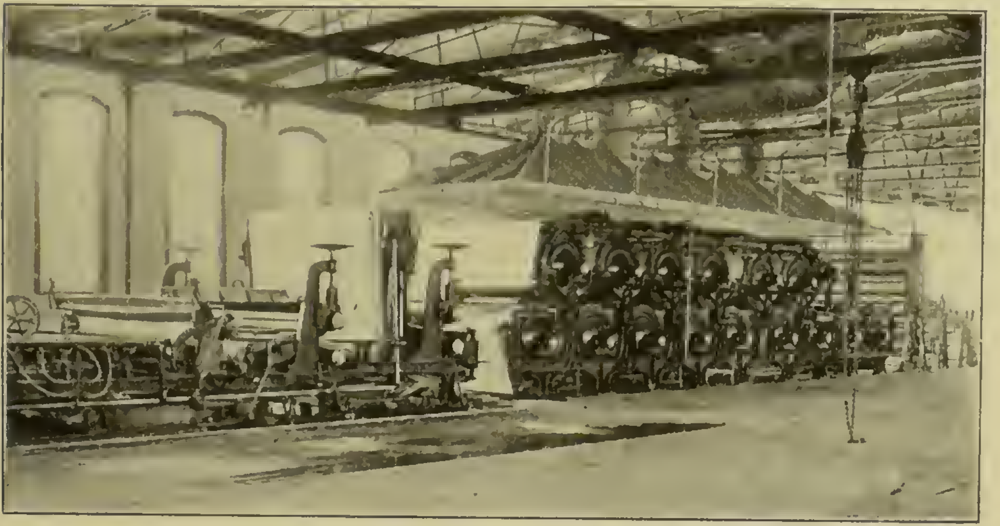
The preparation of wood pulp. — When a log leaves the forest and comes by ox-team, tugging horse-load, or railroad car to the pulp mill, where it is to be turned into wood pulp (the first stage of paper making), any logs which have not been cut in the forest are first cut into sticks of convenient size, and then forced against a rapidly revolving sandstone wheel, over which a jet of water plays. When this alone is done and no more treatment is given, wood pulp is formed, which is run with water on to screens and felts where it is matted into rough loose boards. This wood pulp is used in paper without any treatment chemically, but it cannot be used exclusively for good paper or indeed for paper with practically any strength. Its fibers are too short. It will not mat closely together. The wood treated by chemicals, on the other hand, is just beginning its career when it comes from the cutting machine where it is cut into small pieces. After its cutting it passes through a series of processes on which its strength is to depend. The wood may be treated by the soda process, which is used chiefly for such woods as poplar or basswood; the sulphate process, which is used to some degree for spruce and fir; or the sulphite process. The last is the one most commonly used, and the only one we shall consider here. In this the wood is boiled under pressure, either with sulphurous acid itself or with some sulphite (some compound of the sulphurous acid). This treatment strengthens the fibers of the wood and bleaches them, though a further treatment with chlorine is generally necessary, as the bleaching of sulphurous acids and the sulphites is no more permanent than that of the sulphur dioxide from which they come. The sulphite pulp when finished has long and strong fibers, and forms firm and strong paper. In general the better the paper the greater the percentage of sulphite fibers present in it, the poorer the paper the greater the percentage of wood pulp. Most ordinary papers contain some wood pulp and some sulphite pulp mixed.

Changing wood pulp to paper. — The first process in the change from pulp to paper occurs when the mixture of pulps is thrown into a "beating machine," where the pulp is cut by knives as it whirls through mixed with water. While in the beating machine the pulp is generally thoroughly washed in washers which are placed in the "beaters," and it is here mixed with whatever earthy matter is used for loading, and with whatever dyes are used for coloring. The substance thus formed comes from the beating machine in a finely divided state, and is borne on a sluiceway to the bleaching tanks. In these tanks the pulp is bleached by some chlorine compound, commonly calcium or aluminum hypochlorite. The whitened pulp is mixed with the coloring and weighting matter and then passes to the paper machine.

THE BEATERS

Here the materials which make up the paper are thoroughly ground and mixed.

The paper machine and the making of finished paper. — The mission of the paper machine can be stated in a single sentence. Its work is to remove the excess water from the pulp, felt the paper together and make the fibers mat themselves into a firm closely woven sheet. This is accomplished by sending the wet pulp, with its loosely joined wood fibers and filling, over wire screens where the water drains away, through suction boxes where the air is withdrawn by suction, and over felts where the pulp, as it is dried and compressed, gradually becomes a coherent sheet. To gain a final drying, polish, and surface, this sheet is passed through hot steel rollers or calenders. Thence comes in an endless roll the wide finished sheet.



A Side View of a Paper Machine

Sulphurous acid and its compounds in paper-making. — All these processes depend on one thing, the power that certain chemicals possess to make the fibers of wood available for paper. Sulphurous acid and its compounds probably do more in that direction than any other group of chemicals. Let us study this group briefly.

Preparation of sulphites. — When we ran sulphur dioxide gas into water we saw that sulphurous acid was obtained, hydrogen sulphite. Of the two parts of this acid hydrogen of course gives the general acid quality, leaving the sulphur and three oxgyens of the sulphite as the radical or group of atoms which characterizes this special kind of acid. When we bring metals or their compounds into contact with sulphurous acid, though sulphites do not always form, a number do. The sulphites of sodium and potassium are the best known of the group. They form as in this equation, where acid + base = salt + water:

Sodium hydroxide + hydrogen sulphite = sodium sulphite + hydrogen oxide (water).

Sulphurous acid can form two sets of salts, one of acid salts in which all the hydrogen is not replaced by the metal, and one of normal salts in which all of the hydrogen is replaced by the metal. Thus two sodium sulphites occur, acid sodium sulphite and normal sodium sulphite. Their formulæ and preparation will show their characteristics:

$Na(OH) + H_2SO_3 = NaHSO_3 + H_2O$. $2NaOH + H_2SO_3 = Na_2SO_3 + 2H_2O$.

The test for sulphites. — The strong odor of sulphur dioxide may be observed when an acid is added to a sulphite. The release of this gas under these circumstances is one test for a sulphite. A still better one may be found in the precipitate formed when barium chloride is added to either a solution of a sulphite or to sulphurous acid direct. The white precipitate that falls is barium sulphite. This is a test for sulphites and sulphurous acid:

Barium chloride + hydrogen sulphite = barium sulphite + hydrogen chloride.

$$BaCl_2 + H_2SO_3 = BaSO_3 + 2HCl.$$

It resembles the barium sulphate used as a test for sulphuric acid, but may be differentiated from it by the fact that the *sulphite* is easily soluble in diluted hydrochloric acid, while the *sulphate* is insoluble.

Properties of sulphurous acid. —Sulphurous acid is a powerful reducing agent owing to the ease with which it turns into sulphuric acid by seizure of the oxygen of the air. The acid first reddens litmus and then bleaches it. When exposed to the air sulphurous acid gradually takes up oxygen, forming sulphuric acid. Its strong odor comes from the vaporization of the sulphur dioxide in the solution.

As a whole sulphurous acid serves comparatively few uses, especially in comparison with the acid most clearly related to it, sulphuric acid, whose properties we are just about to consider. Yet it serves one mighty purpose. It makes possible most of the paper we use for the thousand necessities of our daily life.

SULPHURIC ACID

Sulphuric acid in commerce. — Some one has wisely said of this king of all the acids, "The monthly total of sulphuric acid used by a country is a barometer which tells the extent and amount of its manufacturing industries." Few indeed are the finished products which have not somewhere in the process of their manufacture used liquid from the great glass carboys marked, "Sulphuric Acid," "Oil of Vitriol," "Handle with great care." In storage batteries which receive and supply current for trolley car, electric light, and electric automobile; in dyeing and coloring soft chintzes and heavy brocades; in chemical fire extinguishers ready to put out a sudden blaze; in forming grape sugar, glucose; in breaking down phosphate rocks to set free hoarded phosphorus needed for the food of the soil; in electroplating; in a hundred other electrical processes which take advantage of this acid's great power of electrical conduction; in purifying oils by means of its hygroscopic action, its power of removing water; in forming such widely used solvents as hydrochloric and nitric acids; in all these ways sulphuric acid yields to scarce another compound made by man in the diversity of its manufacturing possibilities.

General preparation of sulphuric acid and sulphates. — As we have seen, sulphuric acid may be formed from sulphur trioxide gas, sulphuric anhydride, as sulphurous acid is formed from sulphur dioxide. If we run sulphur trioxide gas into water we get an acid in which, as in all other acids, hydrogen supplies the general acid properties, while the sulphate radical, made up of a sulphur atom and four oxygen atoms, gives the special characteristics. This compound readily dissolves most of the metals forming sulphates and releasing hydrogen, acting in these cases as it did when we made hydrogen.

Zinc + hydrogen sulphate = zinc sulphate + hydrogen.

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

The two processes for manufacturing sulphuric acid. — The manufacture of sulphuric acid is an example of a revolution in methods brought about by scientific progress. There is every reason to believe that the sulphuric acid factory of the future will be a contact factory in which the acid will be produced by catalytic action rather than by the historic English or Chamber process. As these two processes illustrate so well the difference between old and new ways of manufacture, it will be worth our while to consider both methods in some detail.

The contact or catalytic process for making sulphuric acid. — The contact or catalytic or platinum process of making sulphuric acid (for this process is called by all three names), depends for its value on the property possessed by heated platinum of aiding sulphur dioxide gas to take up a part of the oxygen of the air. The hot platinum itself takes no direct part in the reaction, but, by that strange power which we call catalysis, enables the sulphur dioxide to do what it can seldom do otherwise, take up one part of oxygen from the air to form sulphur trioxide. We might express this action graphically thus:

Sulphur dioxide (aided by hot platinum) + oxygen = sulphur trioxide.

$$SO_2 + O = SO_3.$$

Once the sulphur trioxide is obtained the rest is easy. All we have to do is to run the gas into water, and sulphuric acid is obtained.

$$\text{Sulphur trioxide} + \underset{\text{(water)}}{\text{hydrogen oxide}} = \underset{\text{(sulphuric acid)}}{\text{hydrogen sulphate}}.$$

$$SO_3 + H_2O = H_2SO_4.$$

Even if we include every division of the process there are only four parts:

First — Burning sulphur in air, forming sulphur dioxide.

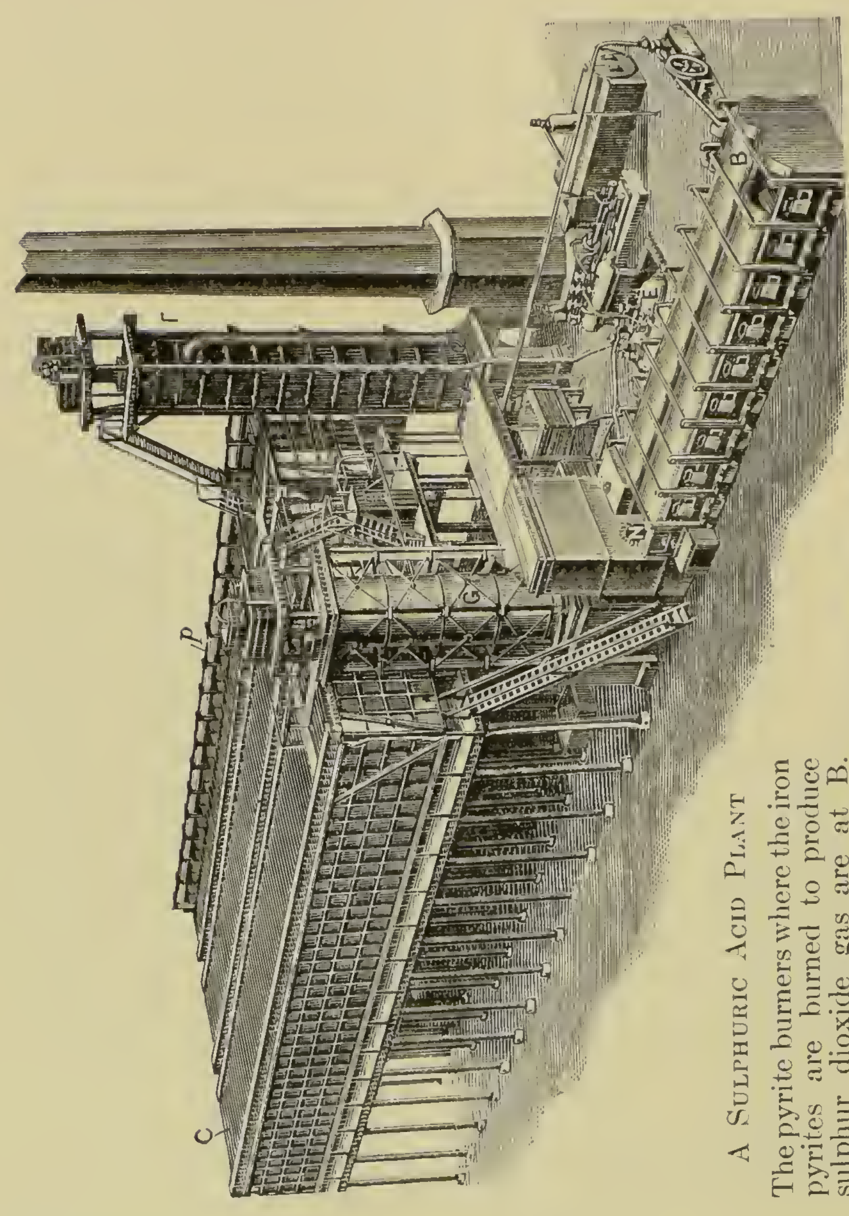
Second — Passing the gas formed over hot platinum, which, acting as a catalytic agent, oxidizes the sulphurous oxide to sulphuric oxide.

Third — Dissolving the sulphur trioxide gas in water and forming the concentrated acid, oil of vitriol.

Fourth — Diluting this acid with water to make ordinary commercial sulphuric acid.

The English or chamber method for making sulphuric acid. — The old process, the English or chamber method for preparing sulphuric acid, can trace its history back to an old monk and alchemist, Basil Valentine, the first man to discover, or at least to record, the existence of the acid. Valentine first made the liquid by distilling iron sulphate, then called green vitriol, a name which has persisted to the present day. He did not stop with that, however, but carried his researches farther, and discovered another way to make the compound, by burning sulphur and niter (sodium nitrate) under glass in the presence of antimony sulphide. He obtained the same acid he had made before, and proved that it could be made from sulphur directly, and not from the comparatively rare and expensive iron sulphate. That discovery of the old alchemist was made more than four centuries ago, but it was not till more than two hundred years had passed that its inherent possibilities were recognized. Then it was rediscovered and improvement followed rapidly upon improvement. First sulphur and niter were heated together in glass. Then, finding glass too expensive, lead chambers for the acid were introduced, and these lead chambers have remained the containers used to the present day, as lead is insoluble in the weak acid, is easily handled, and comparatively inexpensive. To all intents and purposes those two discoveries, the burning of sulphur and niter and the handling of the acid in lead chambers, are at the basis of the chamber process to-day.

The carrying out of the chamber process. — The sulphur used for making sulphur dioxide, which is the first step in any process, is commonly taken from iron pyrites, "fools' gold," which is a sulphide of iron. When this mineral is roasted, the sulphur burns away and its oxide passes off. In consequence a pyrites burner, as it is called, is found in most sulphuric acid plants. The sulphur dioxide gas, whether



A SULPHURIC ACID PLANT

The pyrite burners where the iron pyrites are burned to produce sulphur dioxide gas are at B. The lead chamber into which the oxides pass are at C. The oxides of nitrogen which turn the sulphur dioxide to trioxide are at N. The gases, after passing through the lead chambers, pass up into a tower T (called the Gay Lussac tower) where the oxides of nitrogen left after the action are absorbed and then returned to the chambers through a second tower, the Glover tower G.

obtained from burning sulphur or roasting iron sulphide, is passed into leaden chambers where it is mixed with air, with oxides of nitrogen and with wet steam. The oxides of nitrogen change the sulphur dioxide into sulphur trioxide, which, uniting with steam, forms sulphuric acid. This acid falls and collects on the walls and floors of the chambers. The oxides of nitrogen take oxygen from the air, acting like messengers, and give it to the sulphur dioxide, every molecule of which takes an atom of oxygen and forms sulphur trioxide. The oxides of nitrogen after they have given up their burden rush back for more, and again transfer oxygen from the air to the sulphur dioxide. This exchange continues indefinitely, the sulphur trioxide gas formed being constantly dissolved in water. The resulting liquid is taken off from time to time to be concentrated or diluted to the strength required. The reactions for the preparation of sulphuric acid by the chamber method are somewhat too complicated for use here.

The drying properties of sulphuric acid. — On the various properties which belong to this great acid, many of its uses depend. It eagerly absorbs water from the air or from any water-bearing substance. From such hygroscopic action comes its use as a drier. Like calcium chloride it is used to keep delicate instruments free from the attacking action of the water. Hydrogen, for example, as it comes from the generator is always mixed with some gaseous water. If we bubble the gas through sulphuric acid the acid holds the water and sends the hydrogen on dry. This same hygroscopic action is of great practical importance in the cleansing of petroleum products (such as kerosene and the like) from water. The burning and lubricating power of any oil is of course directly diminished by the presence of water, a substance which will neither burn nor lubricate. Sulphuric acid mixed with the oil under proper conditions cleanses it, purifies it, and can be separated out more readily than can water.

The boiling-point and electrolytic characteristics of sulphuric acid. — The high boiling-point of sulphuric acid makes it available for the production of the most of the other acids. In a mixture, in consequence, such acids pass off as gases while the sulphuric acid remains as a liquid, because these other acids boil at a lower temperature than oil of vitriol. Combined with its other advantages, the power which sulphuric acid, in common with other acids, possesses of conducting electricity provides a market for the thousands of tons annually sold for use in storage and generating batteries and in electroplating baths. As a catalytic agent it assists in making glucose. Its use in preparing soda for the manufacture of glass has already been considered in the chapter on the carbonates.

The mixing of acid and water. The action of sulphuric acid on litmus and on clothes. — Great heat is generated when sulphuric acid and water are mixed together. If water is poured on sulphuric acid the mixture rises to a comparatively high temperature, and may even be raised to the boiling-point when the steam which ensues spatters the acid dangerously. In consequence, when mixing, *sulphuric acid should always be poured on water. Never pour water on this acid.* The heavy acid, sinking into water, is cooled, under such circumstances, below the point of danger. Sulphuric acid is a heavy oily liquid in its concentrated form, is nearly twice as heavy as water, and is strongly acid to every test. When spattered on clothing it turns the cloth red and then rapidly eats it away. The red spot of a sulphuric acid stain should be immediately covered with ammonia. Ammonia neutralizes the acid, forming ammonium sulphate and prevents further action. Great care should be taken never to spill concentrated acid, or, indeed, acid of any strength on the hands or face. Vitriol burns, as they are sometimes called, are serious in the extreme.

Preparation of and test for the sulphates. — The formation of the sulphates, the compounds of sulphuric acid, was exemplified in the reaction by which hydrogen was generated:

Zinc + hydrogen sulphate = zinc sulphate + hydrogen.

Another type of sulphate formation is shown in the test for sulphuric acid with barium chloride. When this salt is added to either sulphuric acid or a sulphate in solution white barium sulphate falls:

Barium chloride + hydrogen sulphate = Barium sulphate + hydrogen chloride.

$$BaCl_2 + H_2SO_4 = BaSO_4 + 2HCl.$$

This white barium *sulphate* is distinguished from white barium *sulphite* by the fact that unlike the *sulphite* it is insoluble in hydrochloric acid. This is the common test for sulphuric acid and the sulphates.

The uses of the sulphates are quite as varied as those of the salts of any acid. Some of them will be considered here and there in our study of the metals.

CHAPTER XXII

PHOSPHORUS AND THE PHOSPHATES

Symbol for phosphorus = P

The uses of a match. — Last night when the dark came down a hundred million lamps sprang into flame. Swaying paper lanterns lighted the smooth streets of Tokio where the jinrikishas race up and down. Torches of pitchy wood gave out a smoky glare where the hunters camped beside the great Amazon. The dark, smoke-stained walls of Swiss chalets, on the edge of that white line from which the snow never fades, threw back the shimmer of home-made tallow candles. Gleaming incandescent gas burners gave brilliant illumination to busy shops and offices, while the study lamp gave out a soft and kindly radiance on the home table. And practically all the lamps of the world, save those the electric current heats, were lighted the same way — with a match.

Small and insignificant as a match may seem, modern man is quite lost without it. His ancestor could use flint and steel, or bear a glowing coal from his neighbor's hearth, when he wished to kindle a fire anew. But to the present generation the match is an absolute necessity, and one so common that its construction is never considered.

Phosphorus and its oxidation. — Phosphorus is the element which gives us the most needed characteristic of a match, and phosphorus is used because this element in any of its forms burns very readily, that is, it easily unites with oxygen to form phosphorus oxides. There are very few substances which oxidize so readily as phosphorus. Leave a bit of the yellow variety out in the air and a thin gray smoke of phosphorus oxide (P_2O_5) begins to rise even at ordinary temperatures, while, if the piece be left for any length of time in the air, the whole of it will burn and change to oxide. In fact, to keep phosphorus in elementary form at all we are obliged to keep it under water and take it out only when it is needed. Even then this element cannot be handled with bare hands since the heat of the hand is sufficient to bring about its rapid oxidation, causing painful and serious burns. Friction ignites yellow phosphorus rapidly. That is seen every time an ordinary match is struck. The lighting of such a match is due to the fact that the heat evolved by the friction is sufficient to bring the phosphorus over the kindling or ignition point. This kindling or ignition point has been defined as that temperature to which a substance must be raised in order that combustion may take place. Passing that point causes phosphorus to burst into flame.

The allotropes of phosphorus. Red and yellow phosphorus. — The yellow variety of phosphorus is only one type of the element. Like oxygen and sulphur this element has an alias, an allotropic modification, to speak technically, which possesses exactly the same chemical structure and which forms the same substances in combination. Nevertheless, these allotropes are wholly different in appearance, one, red phosphorus, being red and opaque or impervious to light, while the other, common phosphorus, is generally of a faint yellow color, and when absolutely pure is almost colorless and quite transparent. Pure yellow phosphorus is no easy thing to obtain, since the element forms an oxide on any provocation, and in consequence soon gets an outside coating of white oxide of phosphorus. Even keeping it under water will not wholly prevent this formation since air bubbles exist in ordinary water and furnish oxygen for oxidation. Red phosphorus, on the other hand, is far more stable. Its brother, the common yellow phosphorus, constantly tends to turn into this form, and when exposed to light grows steadily redder and redder, owing to its partial change into the red phosphorus (known also as amorphous phosphorus, amorphous meaning without definite form). Such a change is almost as if a white man should slowly turn into a red man; the characteristics of red and yellow phosphorus are almost as different as those of the Indian and the Caucasian. Where yellow phosphorus easily takes fire at ordinary temperatures, red phosphorus is practically unaffected and may be left lying exposed in the air for months with scarcely a change in weight or in appearance. Yellow phosphorus cannot be handled with the bare hands. Red phosphorus can be handled with impunity. Yellow phosphorus has a smell like that of garlic or onions. Red phosphorus is odorless. Both allotropes are insoluble in water, but yellow phosphorus dissolves in an organic solvent known as carbon bisulphide, while red phosphorus does not. Yellow phosphorus is extremely poisonous, causes frightful diseases in match factories where workmen are exposed to its fumes, and has produced serious trouble in many cases where matches have been thoughtlessly chewed. Red phosphorus, while by no means an article of diet, has no poisonous properties. Yellow phosphorus has that peculiar property of glowing without visible combustion known as phosphorescence. Red phosphorus possesses this characteristic not at all.

The manufacture of matches. Early types. — Manufacturers of matches take advantage of the different properties possessed by these two allotropes to make different kinds of matches. Back in 1812, the year of the second war between England and the United States, the first matches, "chemical matches," as they were called, came into use. Cumbersome things they were too, whittled sticks which had been dipped in melted sulphur, and with a "head" which was coated with a mixture of sugar and potassium chlorate. When you remember that oxygen is regularly made from heated potassium chlorate, you can see that the reason for putting this substance on the match was to furnish an excess of oxygen to aid its burning. As we shall see in a moment, it is used in match making to-day for the same purpose. Sulphur is used more commonly though, for sulphur burning at a low temperature carries on the fire started by the phosphorus in many of our matches to-day, just as it did in those old "chemical matches" made almost a hundred years ago. It was fifteen years after matches were invented before anybody thought of lighting them by the heat caused by friction. Everybody knew, then as now, that if you rubbed two boards together the friction of the boards made them warm, and knew also that a saw cutting through wood heats both wood and the saw by the friction which ensues. Nevertheless, for fifteen years nobody put that fact and matches together to strike a light. They got a flame instead by dipping their "chemical matches" into a bottle holding asbestos moistened with sulphuric acid. The sugar, the sulphuric acid, and the potassium chlorate when they came together made heat enough to kindle the sulphur, that in time kindled the match, and they had a light without being obliged to strike flint or steel. But think of carrying a bottle holding sulphuric acid around with you every time you wanted to strike a match!

The first friction matches. — With the year 1827 came the first friction or lucifer (the light-bearer) match. This was ignited by rubbing it on sandpaper, and after experimentation it was recognized that such matches gave the best result when they used yellow phosphorus mixed with some oxidizing substance, that is, mixed with some substance which could readily give up its oxygen to phosphorus. Potassium chlorate, niter, lead peroxide, and manganese dioxide all were used for this same purpose, but the chlorates (as they form explosives) have nowadays been practically abandoned for this kind of match.

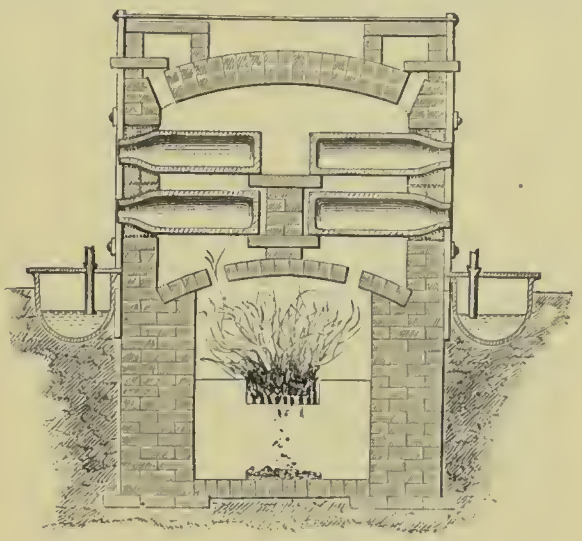
Modern matches. — The whittled stick of 1812 has given place to machines that cut thousands and hundreds of thousands of match sticks an hour from soft pine or spruce. The heads of the sticks cut out for ordinary matches are soaked in different mixtures, of which one that contains sulphur paraffine, phosphorus, and glue is a type. An oxidizing material is mixed with this, and the head is colored some gay color, red or blue, or white and blue. When the match lights, the phosphorus kindles first from the heat of the friction, and unites with the oxygen of the oxidizer aided by the oxygen of the air. The flame of the phosphorus as it turns to its oxide sets fire to the sulphur and paraffine, whose kindling-point is above that of the phosphorus but below that of the wood. The sulphur burning sets fire to the wood and the whole blazes up. It is a whole succession of events each of which depends on the event which immediately precedes it.

Dangers of the common friction match. — Ordinary friction matches are said by eminent authorities to be the cause of thousands of fires. Rats and mice can set them off, children can easily strike them, a footstep on a fallen match may start a destroying conflagration. Safety matches are far better. The chance for accidentally igniting them is much less, and in some countries, especially Switzerland, manufacture of the common friction match containing yellow phosphorus is forbidden.

Safety matches. — Safety matches require two factors for their combustion: the match itself and a preparation on which the match can be struck. The preparation is placed on the side of the match-box and the match is meant to light only on that or on some similar substance. The exact compositions of most safety matches is zealously concealed by the manufacturers as trade secrets, but in many cases the action depends on the fact that red phosphorus can be changed to yellow phosphorus by the application of heat, so that the use of red phosphorus characterizes the safety match. The red phosphorus is commonly placed on the box and then struck by a match which holds oxygen-bearing compounds. The friction of the blow changes a little red phosphorus to yellow phosphorus, which gives just enough heat to bring about combustion on the end of the match. The necessity for bringing together the match and the box before the match will light makes the safety match a great factor in fire prevention.

The preparation of phosphorus. — Phosphorus is probably the only element which is obtained largely from animal sources.

The bones of animals contain considerable quantities of this element, so much of it in fact that most of the phosphorus used to-day comes from the reduction of ground bone-ash with sulphuric acid. The other processes for obtaining phosphorus are two: one by which the phosphates of various minerals are decomposed with production of phosphoric acid (an acid which we shall consider a paragraph or two farther on); the other the direct production of phosphorus from calcium phosphate by the electric current. In both the earlier processes the phosphorus is heated till it turns to a vapor which condenses into a liquid. Solid yellow phosphorus forms from this. The red phosphorus is made by heating the yellow phosphorus for some hours in retorts closed against any possible admission of air which would oxidize it.



MAKING PHOSPHORUS I

Manufacture of phosphorus from bone ash. The charred mixture, left after the ash has been treated with sulphuric acid and with carbon, is distilled in the bottle-shaped retorts shown above the fire. The distilled phosphorus runs down through pipes into troughs of water shown at the right and left. There it condenses.

The reactions for the preparation of phosphorus from the calcium phosphate of bone-ash follow:

1. Action of sulphuric acid producing phosphoric acid.

$$Ca_3(PO_4)_2 + 3H_2SO_4 = 3CaSO_4 + 2H_3PO_4.$$

In this reaction the three calcium atoms of the calcium phosphate join with the three sulphate radicals of the sulphuric acid to form calcium sulphate. The six hydrogen atoms which, in sulphuric acid, join the bivalent sulphate radical as $3H_2$ join with the trivalent phosphate radical, and appear, in consequence, as $2H_3$. The total number of the hydrogen atoms is the same (6), but the way in which they arrange themselves differs.

On heating with carbon the phosphoric acid is converted to metaphosphoric acid, HPO_3.

$$H_3PO_4 = HPO_3 + H_2O.$$

On further heating with carbon, phosphorus is produced.

$$4HPO_3 + 12C = 12CO + 2H_2 + 4P.$$

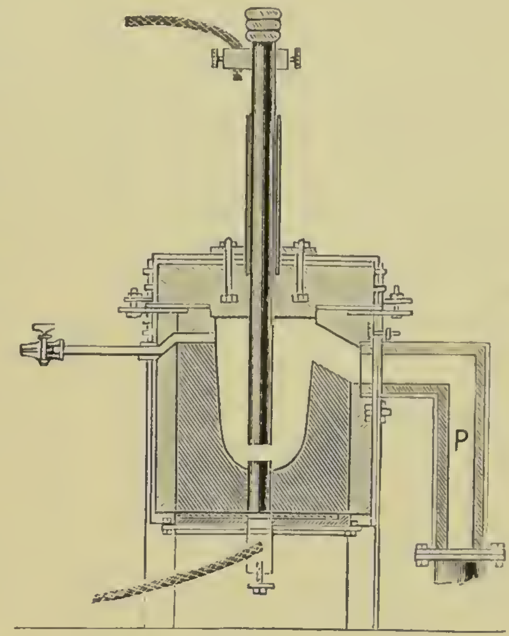
In this reaction there are twelve atoms of oxygen in $4HPO_3$ (4×3). All these atoms unite singly with the twelve carbon atoms to form twelve molecules of carbon monoxide. The hydrogen and phosphorus atoms form elementary molecules.

When calcium phosphate is reduced by the electric current the following reaction takes place:

$$Ca_3(PO_4)_2 + 14C = 3CaC_2 + 2P + 8CO.$$

In this reaction we again see all the eight oxygen atoms in $Ca_3\ (PO_4)_2$ $(4 \times 2 = 8)$ joining with carbon to form eight molecules of carbon monoxide. The three atoms of calcium (3) join with six more atoms of carbon to form calcium carbide and the phosphorus is freed.

Phosphorus in plant and animal life. — Phosphorus is not only an advantage in the way in which it kindles flame, it is an essential to life itself. All living creatures contain a large amount of phosphorus in their bodily construction, and constant additions of that phosphorus are absolutely necessary. This addition of phosphate through the food comes mostly from the various forms of plant life which get their phosphorus from the phosphates of the soil. These phosphates are derived from phosphoric acid, the acid formed by phosphorus, oxygen, and hydrogen. This compound is found in many of the fertilizers and largely, as well, in various minerals, of which one known as apatite may be taken as an example. These minerals, all phosphates of calcium in their main structure, are used very widely for supplying food to the ground. It has been said that no other substance so quickly aids the productiveness of a country as the phosphates that it contains, for without phosphorus or some form of phosphates, animal life would cease to exist. Beside the phosphoric acid, which is made by the action of nitric acid on phosphorus, there is yet another acid, phosphorous acid, which contains one atom less of oxygen, whence its name, the "ous," indicating as before that in the case of sulphurous acid the form is the lower one.



Making Phosphorus II

Manufacture of phosphorus in the electric furnace. In this process the calcium phosphate is mixed with carbon and heated between electrodes similar to the electrodes of an arc light which appear vertical in the picture. The phosphorus distils over through the pipe, and is condensed below.

Phosphate baking powders. — One important use of the phosphates is found in the manufacture of baking powders which are used to the amount of some fifty million pounds a year. Not all of these are phosphate powders, most of them are cream of tartar or alum powders. But all of them act after the same fashion. Some acid principle and some basic principle combined give off carbon dioxide gas. In practically all the powders, baking soda, which is one of the two carbonates of soda, is used. Some acid salt, a phosphate of aluminum for example, is added to this baking soda, whereupon carbon dioxide gas is evolved. The carbon dioxide gas which comes off throughout the mass when the carbonate and the acid principle act one on the other raises the bread and makes it porous. Phosphate powders are made from acid phosphate of lime and generally contain beside this compound and baking soda some small proportion of another substance which serves as a filler to hold the rest. The carbon dioxide produced by baking powders raises the bread, and generally makes its presence known by the pores in the bread where the gas existed. The solids left by phosphate baking powders are likely to be harmless salts of calcium and of sodium. It has often been claimed that phosphates serve to nourish the brain, but no ground for that assumption has ever been found. The brain foods, so called, which contain phosphoric acid as their chief ingredient, have been known to contain it in excess.

CHAPTER XXIII

THE THEORY OF SOLUTIONS AND THE EFFECT OF ELECTRICITY ON CHEMICAL ACTION

The relation between chemistry and electricity. — From the very beginning of organized chemical study the relation between electricity and the reactions which take place between certain groups of chemical compounds has been a source of much ardent investigation. Even the early researches of Volta, of Galvani, and of their followers, which resulted in the production of the galvanic cell, are indissolubly bound up with electro-chemistry, while Faraday, the wizard of electricity, found time among his manifold occupations to open more than one door for the student of this subject, and point out the road beyond.

Electrical generators. — On several occasions we have spoken of the action of electricity on chemical compounds. Let us here collect some of the scattered facts which lie behind us, add some new conceptions and examine the total. First and foremost we know, from the law of the conservation of energy, that electricity cannot be generated spontaneously. This means that electricity must be obtained by the transformation of some other type of energy, though most of us think of the electrical current as being directly produced by one of two types of generators, either by the big dynamos which furnish power for the trolley lines, or else by the dry battery which furnishes sufficient electric power to ring the door-bell or give the spark of the engine of the automobile.

The action of an electrical current in general. — Those two generators are examples of machines which transform energy. The dynamo turned by steam power or by water power is an example of heat or mechanical energy turned into electrical energy. The dry battery is really a wet battery (though its water does not show), and inside of it the chemicals react among themselves in such a way as to enable chemical energy to turn into electrical energy when the circuit (which is, practically speaking, the wires through which the electricity can travel on a continuous path) is closed. To understand the action of such transformers of energy, a word or two must be inserted here concerning the theory of electrical action. No matter in what way we generate our electricity there is always a flow of the current *from* one point *to* another. It must come from a point *more highly* charged to another *less charged*, it must run down, as water flows down hill. Electricity can be considered as passing from a higher to a lower point, or, as we name them for convenience' sake, from a positive terminal or electrode to a negative terminal or electrode. When the electric current jumps from one pole or higher terminal to another pole or lower terminal in the arc light, it is passing from a positive to a negative pole. The current which feeds the arc light leaves a dynamo from a positive pole and comes back to a negative pole. To ring the door-bell a current leaves a battery by a positive electrode, and comes into the battery again through a negative electrode.

CONDUCTORS

A street scene with its many crossing conducting wires, carrying current for trolley cars, telephone and telegraph wires, light and power. These are metallic conductors. In chemistry we chiefly deal with conducting solutions.

The action of an electrical current in a conducting solution. — If we pass an electric current into a solution which conducts electricity, the current goes in at one side at a positive terminal, passes through the solution, and comes out on the other by way of a negative terminal. When an electric current passes through a conducting solution the positive terminal at which it enters is called the anode, the negative terminal by which it leaves is called the cathode. Both the positive and negative terminals are spoken of as electrodes. A word of further explanation should be inserted here concerning these conducting solutions. Under ordinary circumstances electricity does not travel through the air. To go from one point to another on a circuit it must have a conductor through which to pass. Such conductors are likely to be metals like copper or solutions of inorganic compounds like sulphuric acid. Non-conductors are likely to be non-metals like sulphur or organic liquids like kerosene. If we pass a current through a metallic conductor, the chemical condition of the conductor is not apparently changed by the influence of the electricity. If we pass a current through a conducting solution, marked chemical changes occur. In this latter case the conductor is broken down by the passing current, and does not resume its original form. It is change of this type that we are to consider here.

Electrolytes and non-electrolytes. — Solutions which conduct electricity possess the quality of polarity, that is, of differentiation into positive and negative parts, and show this characteristic to a marked degree. That sentence means that solids, dissolved and forming a conducting solution, break down into tiny positive and tiny negative parts. Not all solutions, as has been said, carry electricity. Those which do not, those through which an electric current will not pass, are called non-electrolytes. Such substances are commonly organic compounds like sugar. Solutions which conduct the electrical current and which, while they conduct it, are decomposed by it, are called electrolytes. Such substances are commonly inorganic compounds like common salt.

Speaking a little more exactly, they are apt to be solutions of acids, bases, and salts, or else fused (that is, melted) solids. The breaking down of a conducting solution is known as electrolysis. A commercial use of electrolysis will be discussed in the chapter on sodium.

Differences between electrolytes and non-electrolytes. — We can readily distinguish experimentally between electrolytes and non-electrolytes by making a solution from a member of each class and passing a current through it. If we make a solution of salt a part of a circuit with which is connected some instrument which detects the presence of an electric current, we find a clear swing of a magnetic needle which indicates the passage of the electric current. If we now insert a solution of any other electrolyte in the same solution the needle still shows that the current is passing. If, on the other hand, we insert a solution of sugar or any of the other non-electrolytes in the circuit the current instantly ceases and the needle refuses to move. The sugar solution has stopped the current. It has interposed a wall as real as if the circuit were completely broken by a barrier of stone or air.

The theory of electrolytic dissociation. — Most of the inorganic compounds we consider in this book are electrolytes when in solution. Most of the phenomena which we have observed have probably been due to a considerable degree to the electrolytic properties which these substances possess. To explain such action a Swedish physicist, Arrhenius, brought forward in 1887 a Theory of Electrolytic Dissociation.

This theory involves the explanation of many phenomena, and some of its conclusions rest upon calculations far too abstruse for this work. Nevertheless, from this wide-reaching theory we may be able to gather some points of value closely related to the matter discussed in the preceding chapter on equations.

The difference between electrolytic substances in solution and out of solution. — According to the beliefs of Arrhenius, an electrolyte out of solution is not the same thing chemically as an electrolyte in solution. An electrolyte, like sodium chloride, for example, out of solution is in a comparatively uncharged state. This uncharged state is the common solid state where an atom of sodium is joined with an atom of chlorine to form a molecule of sodium chloride. Dissolve sodium chloride in water and part of its molecules break down into atoms which are in a new condition. The breaking down of an electrolyte in solution is called dissociation, and the molecule under such circumstances is said to dissociate. Every molecule when dissociated breaks down, not into ordinary sodium and chlorine atoms, but into sodium atoms possessing a charge of positive electricity, and into chlorine atoms possessing a charge of negative electricity. Bodies such as these, as we have said, are called ions; and every electrolyte in solution dissociates into ions to a greater or less degree. If we had the power of vision possessed by the Dæmon whom we conceived, and could look into a glass of water in which salt has been stirred, we should expect to see that it contained a number of sodium ions (atoms bearing a positive or plus charge of electricity) and an equal number of chlorine ions (atoms bearing a negative or minus charge of electricity). In such a solution there would be also present a number of undissociated molecules of sodium chloride. There would be a constant interchange between the dissociated part of the solution (the ionic part) and the undissociated part (the molecular part). Some of the ionized bodies would be constantly becoming undissociated molecules. Some of the undissociated molecules would be constantly becoming ionized. But despite this interchange the percentage of ionized parts to non-ionized parts would always be constant. All electrolytes ionize or break up into ions to greater or less extent. Now to repeat our definition of an ion, we may define it as an atom or a radical (a group of atoms acting as a single atom) which possesses a charge of electricity. As regards the production of these ions we may quote from Talbot and Blanchard's "Electrolytic Dissociation Theory," page 12:

"Ions may be produced by the dissociation of larger electrically neutral molecules (that is, of the molecules of compounds as ordinarily expressed by their chemical formulas) into smaller ones bearing, respectively, *equivalent* amounts of positive and negative electricity. In every solution the aggregate (the total sum) of charges on the positive ions which it contains must be exactly equalized by the sum of the charges on the negative ions present, otherwise the solution as a whole could not be electrically neutral."

How sodium chloride acts in solution. — Let us view some parts of that question in the light of the solution of sodium chloride already mentioned. The molecules of sodium chloride are electrically neutral. No part of the molecule exhibits any positive or negative tendencies towards external things, though it is entirely probable that there may be an internal electrical polarity. These molecules break up into smaller parts or divisions which are sodium ions bearing positive charges of electricity and chlorine ions bearing negative charges of electricity. There must be just as many positive charges on the sodium ions in the solution as there are negative charges on the chlorine ions, because there has been nothing added or taken away, and the solution was in electrical equilibrium when we started. That equilibrium meant that the number of positive charges were exactly balanced by the number of negative ones.

The relation between valence and ionization. — Right in here comes the bearing of this theory on valence, which is the balancing of elements in combination. Viewed by this light the answer to many questions of solution reaction becomes clear. Why does one atom of chlorine unite with one atom of silver to form a molecule of silver chloride when two atoms unite with copper to form copper chloride and three atoms unite with iron to form ferric chloride? The answer in terms of ions would be as follows: First, a silver ion carries one positive charge while the chlorine ion carries one negative charge. One positive and one negative produce equilibrium.

$$Ag^{+} \text{ added to } Cl^{-} = AgCl.$$

Second, every copper ion carries two positive charges to the chlorine ion's one. Two chlorine atoms are required to form two chlorine ions which will balance the copper. The copper ion can be written with two plus signs to indicate two charges, Cu^{++}.

$$Cu^{++} + Cl^{-} + Cl^{-} = CuCl_2.$$

Third, iron in the ferric state forms an ion which carries three positive charges. Three chlorine ions are required to come into equilibrium and form the compound,

$$Fe^{+++} + Cl^{-} + Cl^{-} + Cl^{-} = FeCl_3.$$

We do not know just why the silver ion carries one charge, the copper ion two, and ferric ion three, but we can be reasonably sure that they do.

Further consideration of valence. — Almost every reaction given in the inorganic portion of the preceding chapters becomes simple when this theory is applied to it. In solution, reactions of any (*acid* + *base* = *salt* + *water*) type become ionic. This is exemplified in the following examples:

(*a*) $Na^{+}(OH)^{-} + H^{+}Cl^{-} = Na^{+}Cl^{-} + H_2O.$
$(Na^{+} + OH^{-} + H^{+} + Cl^{-} = Na^{+} + Cl^{-} + H_2O.)$
(*b*) $2Na^{++}(OH)^{=} + H_2^{++}(SO_4)^{=} = Na_2^{++}(SO_4)^{--} + 2H_2O.$
$\left\{\begin{matrix} Na^{+} & (OH)^{-} \\ Na^{+} & (OH)^{-} \end{matrix}\right\}$
(*c*) $K^{+}(OH)^{-} + H^{+}(NO_3)^{-} = K^{+}(NO_3)^{-} + H_2O.$

In none of these cases is the water dissociated perceptibly. The other products, NaCl, Na_2SO_4, and KNO_3, dissociate in proportion to the dilution of the solution. In general the more dilute the solution the more these substances ionize. For all practical purposes, however, we consider that strong electrolytes dissociate completely and treat the reaction in which they occur accordingly.

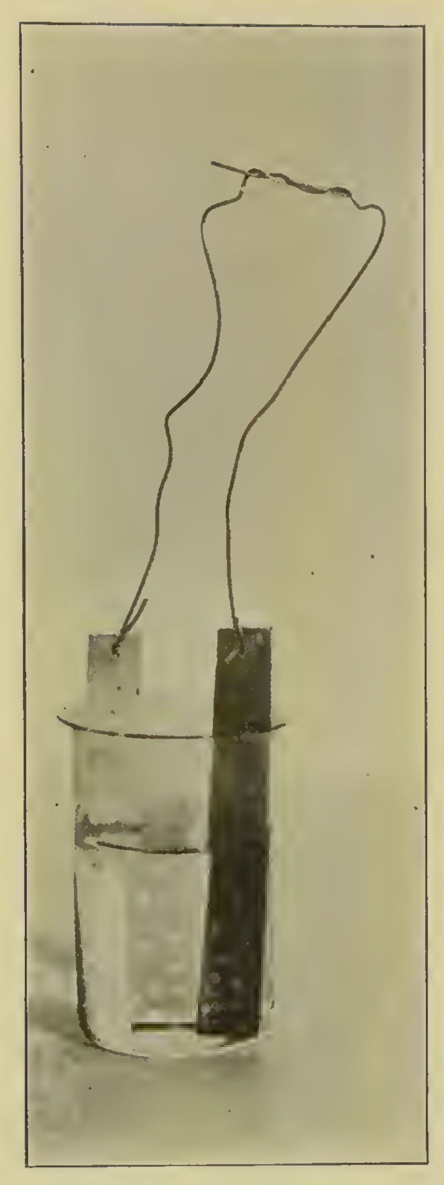
In reaction (*a*) one product of the reaction is $Na^{+}Cl^{-}$, an ionized substance, sodium chloride, which is in solution and is therefore dissociated. The other product of the reaction is water, H·OH or H_2O, which dissociates so slightly as to make its ionic existence hardly worth noticing.

Reaction (*b*) illustrates the necessity for using two molecules of sodium hydroxide in combination with sulphuric acid. The sulphate ion (SO_4) carries two negative charges in its ionic state. It requires two positive $(Na)^+$ ions to bring it into equilibrium. The two positive $(H)^+$ ions of the acid carry two positive charges. They require two negative hydroxyl ions to form two molecules of water. In any similar reaction a sufficient number of molecules, two, three, four, or more, must be taken to bring every part of the equation into equilibrium.

The writing of ionic and molecular symbols. — Another point concerning the equations given above should be carefully noted. The positive and negative signs + and − indicate an ionic condition. Symbols without these signs indicate a molecular condition. The symbol for nitric acid given in the above reaction (*c*) is written with + and − signs to indicate that it is dissociated into ions. The symbol of water is not written with the + and − signs, because of the fact that water is so slightly ionized. The reaction last given indicates, therefore, that the potassium hydroxide, the nitric acid, and the potassium nitrate are all in an ionized state, while the water is in the molecular state. Ionic signs are generally written at the upper right hand corner of the symbol.

The electrical charges carried by the ions. — Valence as considered in the chapter on equations, and valence as considered here, are closely related. This word has already been defined as the power which an element possesses of replacing or combining with hydrogen. Chlorine, for example, was said to have a valence of one, that is, to be monovalent, because one atom of this element combined with one atom of hydrogen. The reason for this in the case of electrolytes may be found in the condition of the hydrogen ion. This ion carries a single charge of positive electricity. Any ion which carries a single charge of negative electricity will form a state of neutral equilibrium with the hydrogen ion. Any such ion would be monovalent and the atom from which the ion was formed would be monovalent also. On the other hand, any ion which exactly replaces a single ion of hydrogen must carry an equal charge of positive electricity. Such an ion is monovalent. Any ion which combines with or replaces two ions of hydrogen must carry two charges of negative or two charges of positive electricity. Such an ion is bivalent. It has a valence of two. Two of its atoms combine with or replace two atoms of hydrogen. In like fashion a trivalent element or one which has a valence of three must carry three charges of positive or three charges of negative electricity to enable it to combine with or replace three atoms of hydrogen. The valence of atoms in compounds that are not electrolytes may not be definitely settled by this theory, but the whole relation between valence and the number of electrical charges that an atom is capable of carrying appears to be close. Once more let us say that a careful study of the reactions in the text with constant references to the tables of valence given in the chapter on equations will prove of great value.

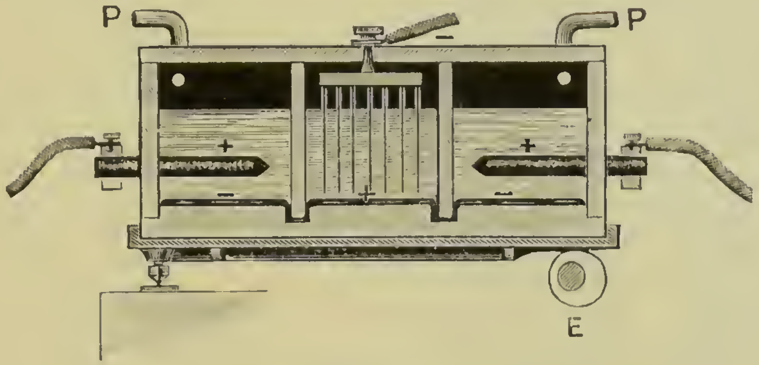
The electrolytic action of a single fluid cell. — The commercial use of electricity in chemical processes, yearly becoming more wide-spread, is closely interwoven with the theoretical developments of this subject. In the single fluid cell, if we disregard the action of the films of hydrogen formed on the electrodes, we may find an example of the workings of electrolysis. A single fluid galvanic cell generates electricity by turning chemical energy into electrical energy. Its positive pole, from which the current flows within the cell, or anode, is made of zinc. Its negative pole, or cathode, is made of copper. These electrodes stand in a solution of dilute sulphuric acid, which is an electrolyte. When the electrical circuit is closed the current passing from the zinc to the copper within the cell decomposes sulphuric acid, H_2SO_4, into its ions, hydrogen H_2^{++} and the sulphate ion $(SO_4)^{=}$. The hydrogen is positive, and in consequence its ions fly to the negative pole, the copper. The sulphate ion is negative, and in consequence its ions fly to the positive pole, the zinc. All the time the current passes, this action goes on. Negative ions rush to the anode or positive pole, and the positive ions hasten to the cathode or negative pole. The negative ions which go to the positive pole are called the anions because they seek the anode. The positive ions which go to the negative pole are called cations, because they seek the cathode.



GENERATING ELECTRICITY
A single fluid galvanic cell. The current passes from the zinc on the left to the copper on the right through the liquid, sulphuric acid, which is broken down by electrolysis. The current then passes from the copper through the wires and back to the zinc outside the cell. Hydrogen bubbles evolved at the zinc electrode are passing across to the copper electrode.

When anions or cations reach the electrodes to which they are bound they discharge the electricity which they carry, cease to be ions, and appear as atoms, molecules, gases, or metallic films according to their nature.

Electroplating in general. — The electrolysis of copper sulphate illustrates a simple type of electroplating. When an electric current is passed by way of platinum electrodes through a copper sulphate solution the ions present in the solution are Cu^{++} and $(SO_4)^{--}$. The copper is the cation, the sulphate ion the anion. When copper sulphate dissociates the positive copper ions go to the negative pole. When these cations reach the cathode they lose their charge and deposit on the platinum electrode as metallic copper. This action copperplates the electrode. The anion $(SO_4)^{--}$ loses its charge when it reaches the anode, and oxygen (due to a reaction between the sulphate ion and water) is evolved.



An Apparatus for Decomposing Solutions by Electrolysis

This is not an electroplating apparatus, but it illustrates the way in which electrolytic action takes place.

Silver plating. — In silver plating we find a slightly more complicated case of electrolysis which we may try to reduce to simple general terms. The purpose of silver or gold plating is to cover a low-cost metal, or an alloy of baser metals, with a coating of precious metal. The spoon or fork which is to be plated is first cast in the form desired. It is then made one of a whole series of spoons or forks which constitute the cathodes of an electric battery containing silver cyanide dissolved in potassium cyanide. The anodes are silver plates, immersed on the other side of the solution. When the current passes through, the silver cyanide ionizes something as the copper sulphate did, in the example just cited, and the positive silver ions fly to the negative poles, in this case to the spoons, which form the cathodes. When they reach the spoons they lose their electric charge and deposit as metallic silver over the whole submerged surface of the spoons. Cyanogen, the anion, goes to the metallic silver which forms the anode, where it combines to form fresh silver cyanide. This replaces the silver cyanide that was lost and keeps the supply of silver in solution constant.

Faraday's law. — When we come to examine the amount of silver deposited on a spoon in a silver-plating bath, or the amount of copper deposited on a platinum electrode, we find a law which governs the deposition of the metals. We owe the discovery of this law to a man whose name appears often in any work on chemistry, Michael Faraday. Faraday discovered that on passing a fixed quantity of electricity through different electrolytes the ratio between the quantities of separated matter obtained by electrolysis was the same as that which existed between their chemical equivalents. To illustrate this by a standard example, suppose we placed solutions of three electrolytes, hydrochloric acid, sulphuric acid, and silver nitrate, in a circuit through which passes a fixed amount of electricity. The products of the electrolytic dissociation of these three compounds would be, first, hydrogen and chlorine; second, hydrogen and oxygen (oxygen being the gas liberated when the anion, SO_4, reaches an anode); and third, silver and oxygen. If we weigh the gases and the solid silver obtained we shall find that the ratio between them is as follows; if hydrogen is taken as 1, then chlorine will be 35.5, oxygen will be 8, and silver will be 108. A reference to the atomic weights of these elements given in the table shows at once that hydrogen, chlorine, and silver have been liberated in the ratio of their

MICHAEL FARADAY. (1791–1867)
A great electrician and electrochemist, who opened the way to much nineteenth-century development

atomic weights. At first glance oxygen seems to have been liberated to only one half its atomic weight 8, instead of 16. To explain the action of oxygen, suppose we insert another electrolyte in the shape of a solution of copper sulphate into the same circuit and weigh the copper deposited in an equal length of time. On doing this we find that the copper deposited 31.75 grams in the same time that it took to deposit 1 gram of hydrogen and 8 grams of oxygen. The copper is deposited in an amount equal to one half its atomic weight, 31.75 instead of 63.5. Evidently there must be some reason for this. The explanation is found in the fact that a substance whose valence is two deposits one half its atomic weight in an electrolytic dissociation. The valence of copper in such cases is 2, and the valence of oxygen is 2. The copper has carried two charges and the bivalent oxygen in this case where the sulphate ion is present is a part of an anion carrying two negative charges. A substance whose valence is 3 deposits one third of its atomic weight in an electrolytic dissociation. In consequence we may use the phrase "chemical equivalent" for this action, and say that the chemical equivalent of an element is its atomic weight divided by its valence. By this law we can tell commercially just the amount of any metal which the passage of any fixed current will deposit. This knowledge is of great importance in manufacture.

CHAPTER XXIV

CARBON

Symbol for carbon = C

The varying forms of elementary carbon. — Wrenched from the drift soil of South African mines to shine in ring, bracelet, tiara, comes the brilliant diamond. From mines in bleak Siberia, warm Ceylon, and blossoming California comes graphite, the black lead of your lead pencil, the lubricant which lessens friction on your bicycle chain, the protector which shields iron work from rust, the covering which conducts the electric current in electrotyping. From mounds heaped high, where smoldering wood is covered from oxygen-laden air by brushwood and by turf, comes the black charcoal which brings to greater heat the kitchen fire. From the high-turned wick of the forgotten lamp comes the sooty cloud of lampblack settling on wall and book and floor, — used when collected for printer's ink and black paint. From bones burned out of reach of air comes bone black, the animal charcoal that furnishes filters to sugar makers. From coal distilled in gas houses come two common types, the gas carbon, which the electrician places in the illuminating lamp of the electric arc, and the coke that feeds coke ovens, furnaces, and railroad engines. From layers of old vegetation long since dead, whose hard-pressed mold and leaves turned black many a century ago, comes coal, the greatest fuel of the world to-day.

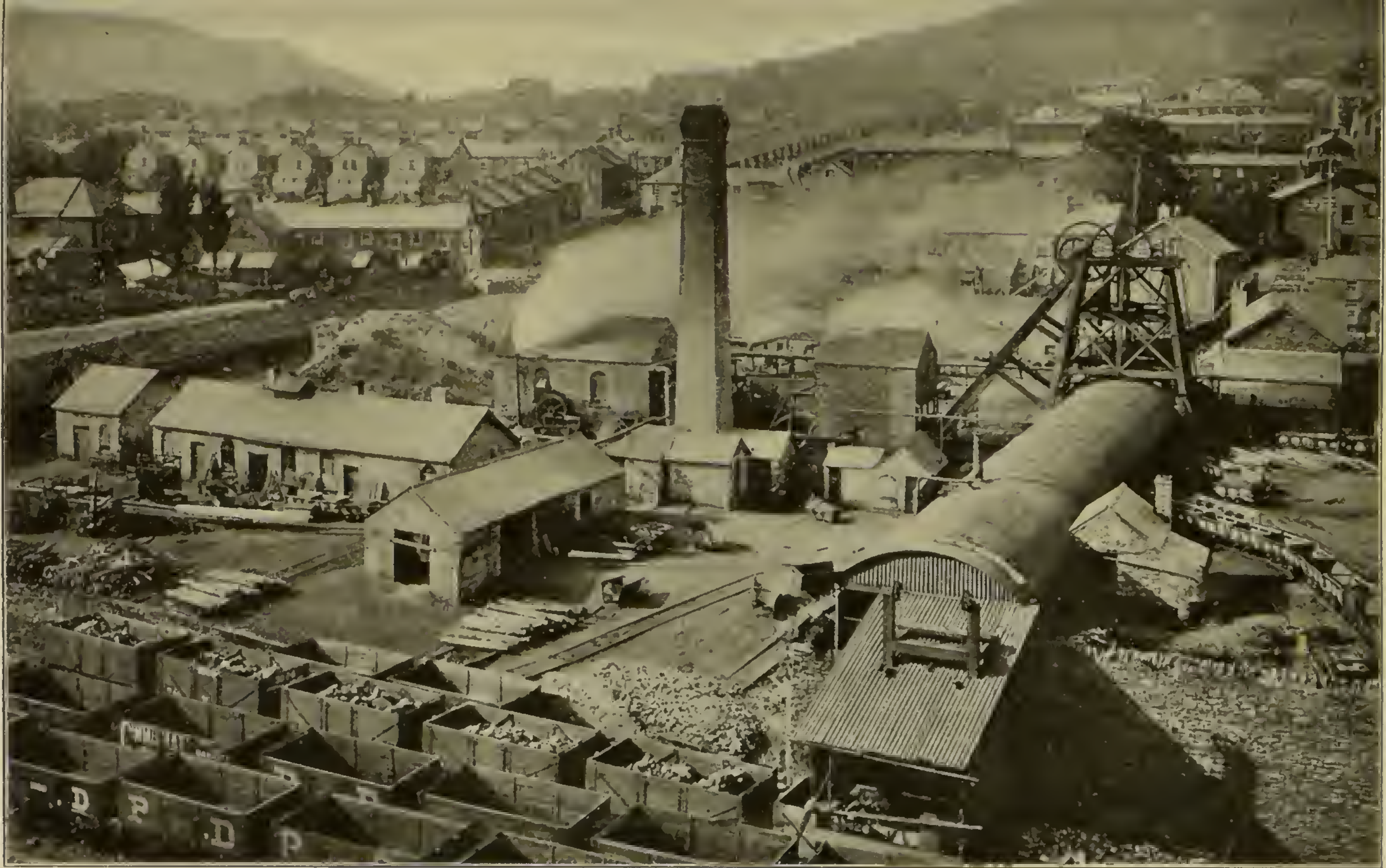
The allotropes of carbon. — Every one of the substances mentioned in the first paragraph is chiefly carbon. Among them three distinct allotropes appear — three aliases or material synonyms of carbon. Two of these three were mentioned first — the diamond and graphite. The diamond is pure carbon. When it burns in air or oxygen it fades into a vapor, which is carbon dioxide and carbon dioxide alone, an action which proves conclusively that the so-called stone is carbon. Graphite, like the diamond, forms carbon dioxide when it burns, but this allotrope leaves behind some traces of the soil. It is an earthy carbon, and like soil contains silica, alumina, or iron. Save coal, all the other types of carbon, charcoal, lampblack, coke, and the like, are grouped under

Diamonds
Kimberly Diamond Mines, South Africa.

the one head of amorphous carbon. Coal in its different forms is generally considered in a class by itself, and is held to be a type of impure carbon.

The diamond. — The diamond has been the precious stone of every age since the dawn of civilization. Eastern rajahs and princes have made war for single gems. Every great monarch wears the glittering jewels in his crown. Every great lady counts her diamonds among her most precious possessions. Found mainly in South Africa, smaller deposits of this jewel have been found in South America, in Australia, and even in some parts of the United States. When shot to the earth through interstellar space meteorites have borne us diamonds from the stars. Yet not all sources from which they come have been able to definitely solve the question of how the diamond originated. The best we can do towards answering that puzzle is to tell the extraordinary way in which diamonds have been artificially made, for the manufacture of diamonds has been proven a possible though by no means a profitable process.

Henri Moissan and the artificial diamond. — To Henri Moissan, the man who isolated fluorine, belongs the credit of originating the brilliant process which resulted in the manufacture of the artificial diamond. The great French chemist knew, as did the rest of the world, that charcoal and the diamond were essentially the same substance. They were both carbon. He believed that the transformation of one into the other might be effected if the proper conditions of temperature and pressure could be obtained. Fortunately, the electric furnace, a piece of apparatus brought to completion by himself, was at his command. By this Moissan was able to obtain the highest laboratory temperatures ever known. That furnace will be discussed in detail later. It is sufficient for our purposes here to say that it imparts to bodies within its crater the terrific temperature of some 3500 degrees C., over 6000 degrees F. In this furnace was placed a crucible containing two things: pure carbon, in the form of charcoal, and iron. The iron



COAL

An anthracite coal mine or colliery. Surface view.

rapidly fused in the intense heat, and changed to liquid form; the carbon melted into the iron. At its very highest temperature the hot mass was suddenly plunged into water. As a result the crust of the fused mixture was suddenly cooled while the interior was red hot. In consequence the cool crust contracting met the hot interior expanding. By this means a terrific pressure was produced, and in some way this enormous compression crystallized the dissolved carbon and formed diamonds. The jewels found were small and of various colors: only a few were white, so the cost of the process was much more than the value of the product obtained. But Henri Moissan had done a great thing. One more substance had been formed in the laboratory by man which had never been made before outside of Nature's great primeval laboratories.

Diamonds

Rough diamonds from Kimberly, showing formation

The colors of the diamond. — The fact that diamonds of various colors were found in the electric furnace process was nothing strange. While we see white diamonds almost exclusively, the precious stones are found in many colors, black, blue, pink, red, yellow, and gray. The white are commonly of the greatest value as jewels, but the other forms are of much value in pointing drills, glass-cutters, and rock borers.

The crystalline form of the diamond. — The diamond is crystalline in form. Its hardness and its wonderful properties of transmitting and reflecting light give it its great value. It splits down from one crystalline form into others, and will not support sudden blows. No precious stone has brought more romance, more hate, more envy, and more desire to man than this.

Graphite. — Diamonds must always be confined to the few, but the second allotrope of carbon is almost literally in every one's hands the civilized world over. It is the lead of your pencil, the every-day pencil that checks goods and writes letters or exercises. The form of carbon in that writing instrument is graphite, commonly called black lead, plumbago, from its resemblance to metallic lead.

Various uses of graphite. — Diamonds could probably cease to exist with but a fraction the loss to the human race that would occur if graphite were to disappear, for this allotrope not only gives the daily service of the pencil, but serves many other common and useful offices. It acts as a protective coating for covering structures with a permanent and durable pigment, which keeps away oxygen and prevents rust. It lubricates machinery as every owner of a chain bicycle knows, for the graphite which keeps his chain running smoothly is this same form of carbon. It is used in manufacturing some kinds of gunpowder, and it is of great value in the preparation of crucibles, and like vessels, which are to resist high temperatures and great extremes of heat and cold. Unlike the diamond, graphite conducts electricity, and is used in electrotyping.

The manufacture of lead pencils. — The lead pencil, while dependent on graphite for the mark which it leaves, is not all carbon. In its manufacture clean graphite is mixed with clean clay, forming a pasty mass which is forced by great hydraulic pressure through tin plates perforated with small holes the size of the black lead rod. The mixture comes through in the form of the lead of an ordinary pencil, and is finally incased in a wooden covering.

Properties of graphite. — Graphite, like the diamond, is proved to be carbon by its formation of carbon dioxide when burned at a high heat. It was formerly found in large quantities in England; the sources of supply to-day are Siberia, the Orient, New York, and California. It is a dull black in color, is very reluctant to combine with any other substance, and extremely permanent.

Amorphous carbon. — Though graphite is commonly noncrystalline in structure, yet it is sharply differentiated from the third allotropic class of carbon forms, which are grouped together under the general heading of amorphous carbon. In this division may be found lampblack, the soot which pours off in a steady stream when you turn the wick of your kerosene lamp too high, the wood charcoal which feeds the tiny fire of the "hot chestnut" man on the first stinging days of fall, and the animal charcoal that whitens your sugar. While none of these are pure carbon, all are made up largely of this great element, and all give carbon dioxide as the chief product of their combination.

Lampblack. — As we have said, lampblack rising from a flame is uncombined carbon, that part of the flame which does not oxidize and turn to colorless gases. Black and finely divided, it makes an excellent base for black paints or pigments, for printing inks, and for the blacking that polishes stove or shoe.

Charcoal. — The charcoal burners' hut and mound have been familiar scenes for centuries in the great forests of Europe, though wood charcoal, uncombined carbon as it is, is more difficult to obtain than is lampblack. The wood used for preparing charcoal is cut into billets of moderate size and stacked together something as a quantity of huge pencils might be grouped in a bundle. This heap of sticks is covered with turf and burned out of contact with the air, thereby preventing the change of much of the carbon to carbon dioxide, but allowing a number of chemical compounds to be formed which are not produced when wood is burned in the presence of oxygen. While various modern methods, such as heating wood in closed iron retorts, have come into use, all are based on the same general principle. Such a process as this of charcoal burning is known as destructive distillation, and deserves a word here.

Destructive distillation. — As we have seen, oxidizable substances, when heated in the air, burn, forming oxides. When such substances are heated in closed vessels, as wood is heated in charcoal burning under a close covering of turf and soil, the cutting off of the supply of oxygen prevents any large amount of oxidation. Instead of producing oxides, complex substances, such as wood or coal, generally break down under such circumstances into other simpler substances made from the elements which compose them. In consequence, destructive distillation has been defined as "the resolution of a complex substance into simpler vapors and gases under the influence of heat out of contact with the air." Coke is the main solid left after such distillation of coal, charcoal after such distillation of wood. Some of the varied substances produced in this way will be considered later.

Uses of wood charcoal.— As wood charcoal contains a greater amount of available fuel, volume for volume, than wood, it gives off a correspondingly greater heat when it is brought to a higher point of ignition; therefore its services in giving a hot fire. Carbon in its pure state is very inert, very slow to combine at the ordinary temperature of the air. In consequence, wood posts which are to be set into the soil are charred to form a preservative coating of charcoal, where they are to be subjected to the destructive influences of the ground.

Charcoal as a deodorizer. — Charcoal is a great absorber. From this property comes its use as a deodorizer. Foul gases rising from sewers are absorbed by charcoal filters placed in manholes. Liquids filtered through charcoal leave many of their impurities behind, though the ordinary domestic filter, whether filled with charcoal or other filtering material, if unchanged, is more likely to be a source of disease than of health. Once it has absorbed its full capacity of solid matter, the water passing through it thereafter is passing through a bed of impurities. When household filters are used, regular cleansing and changing of their contents is a necessity.

The filtering power of animal charcoal. — The filtering power of wood charcoal is far less than that of animal charcoal, the finely divided carbon which comes from the destructive distillation of bones. This is an extremely valuable filtering material, and because of its power of removing coloring matters it is used extensively in sugar manufacture. The white sugar of daily use owes much of its fine color to the filtration which it received while passing through what are known as charfilters, huge filters containing this finely divided charcoal. Great quantities of charcoal are used in sugar refining.

The density of charcoal. — Ordinary charcoal, though commonly half as heavy again as water, floats in water because of the gases held or included within the pores. When these gases are removed by heating, charcoal sinks in water. This form of carbon is a bad conductor of electricity. The various forms of charcoal made from different woods burn at various temperatures.

Coal. — The black coal of the coal-bin is an impure form of carbon, which, in turning to carbon dioxide, heats the house or cooks the dinner, drives the locomotive or turns the mill wheel. Roughly speaking, coal generally comes under one of two headings. It is either soft, bituminous, or it is hard, anthracite. The difference between the two does not depend upon appearance and feeling alone. Anthracite coal contains a larger amount of carbon than bituminous coal. Lignite, a fuel commonly considered under this general heading, is a brown coal which contains less carbon than bituminous, and

COAL

Bringing a train of loaded coal cars from a bituminous coal mine.

peat, a fuel cut from the peat bogs, contains the least carbon of any of the group.

The stages through which coal passes during its formation. — The varying amount of carbon present in peat, lignite, bituminous, and anthracite indicates clearly the various stages through which our fuel passes. Thousands of years ago, regions such as Pennsylvania and Wales, where we find coal to-day, were huge semi-submerged places like marsh lands, where bodies made up of carbon and oxygen, such as the dying tree, and the luxuriant decaying vegetation of swamp and bog, finally, by some peculiar decomposition, lost much of the hydrogen and oxygen, and changed in large part to a bed of carbon. In its first stage such a coal region was like a peat bog; its percentage of carbon, and, in consequence, its fuel value, was low. Gradually, as the centuries rolled on, harder and harder grew the peat. In its next step the fuel lost still more of the other ingredients present besides carbon. A coal resembling lignite formed. As the process continued, a true coal, bituminous, mainly made up of carbon, was produced, and finally after long ages came anthracite, the coal which in its purest state has lost almost every ingredient but carbon. Compression and heat probably played no small part in the process. Many a natural chemical process must have taken place in the change, but the coal we have to-day came from the forests and swamps which existed in the dawn of life on

COAL

A mile under ground in a colliery.

this planet, as is shown by the delicate tracery of leaf and fern still to be observed in coal deposits. Just how many millions of years ago these deposits began to form no man can tell.

The cyanides. — Before passing to the organic compounds of carbon a brief word should be inserted here concerning one of the inorganic compounds of this element which appears in numerous forms. The cyanides, a group which possesses the baleful characteristic of being the most deadly of all the inorganic salts, are characterized by a radical which contains carbon and nitrogen in equal parts, one atom of carbon and one of nitrogen, and whose name is cyanogen. The radical of cyanogen = (CN).

Prussic acid. — Prussic acid, hydrocyanic acid, or hydrogen cyanide (HCN), is the acid from which the cyanide salts are obtained. The gas and its solution are a deadly poison even in small quantities. Laurel leaves, bitter almonds, and the stones of several fruits yield this acid when distilled.

Potassium cyanide. — Potassium cyanide (KCN), the most important salt of hydrocyanic acid, is almost as poisonous as prussic acid. It is of much greater practical importance than the acid, as it is of value in photography, in the reduction of metals and as a laboratory reagent. The cyanides of iron are used in analytical chemistry, but find their greatest use in the production of Prussian blue, one of the strongly established mineral colors, whose use as a dye was wide-spread fifty years ago before aniline colors appeared. The red prussiate of potash (potassium ferricyanide, $K_3Fe\,(CN)_6$), another of the cyanides, is made up of potassium, iron, and cyanogen. It is used with indigo in obtaining white patterns on a blue ground. Some of the extremely explosive compounds known as the fulminates contain this radical cyanogen, and owe to it much of their explosive qualities. The yellow prussiate of potash, (potassium ferrocyanide) = $K_4Fe(CN)_6$, belongs in this group.

CHAPTER XXV

ORGANIC CARBON

The protean forms of carbon. — To the Greeks Proteus was the divinity of changing form. To him belonged the power of assuming any character, any form at will. Pursued by an enemy, the god turned to a small gray rabbit, scampering through the rustling leaves, to a leaf whipped in the breeze, to a still pool mirroring the azure of the sky, or to a wandering breeze stirring the crimson poppies of some southern sloping hill. In each and every form he was Proteus. In each the god himself existed unchanged, no matter what the outer shape. In the following chapters we are to study carbon in its relation to organic chemical compounds. We shall find it in thousands of forms, in foods and in clothes, in aromatic gases and in inflammable liquids, but it will always be carbon. We may say, indeed, that of all the chemical elements carbon is most protean, most like Proteus. Found far and wide, carbon, in each and every form, is a determining factor in molding the characteristics of the compound.

Organic chemistry. — Carbon is present in almost every substance which comes from or depends on nature's laboratories of the soil. It is found in the living green of the earth and in the animals which consume the plenty of the harvests. Indeed, the whole science of organic chemistry is now defined as the study of the compounds of carbon. And organic chemistry long since surpassed in extent the field of inorganic chemistry. Here follows a brief comparison between the study of the carbon compounds (organic chemistry) and the study of the metals, the non-metals, and their compounds (inorganic chemistry).

The old untrue definition of organic chemistry. — First of all, it must be noted that there is no fundamental difference between the two divisions. Organic and inorganic substances alike obey the great chemical laws of constant, multiple, and reciprocal proportions, the general laws of conservation of matter and of energy, and the gas laws of Gay Lussac, of Charles and of Boyle. In years gone by it was wrongly supposed that there was a difference. It was believed that organic substances, like the wheat that grew in the northwest, or the sugar-cane that came from Cuba, like the bread from the wheat flour, or the alcohol from the molasses of the crushed sugar, each and all of these could be obtained only when a mysterious "vital force" was present. The phrase, "vital force," meant that some life process played a part in the action when the substance was formed. Organic chemistry was then mistakenly defined as the chemistry of those compounds which came into existence through the agency of animal or vegetable life; so the starch of the potato and the sugar of the sugar-cane were cited as excellent organic types produced in this way. Inorganic chemistry was defined as the study which related to those elements and compounds such as iron, oxygen, water and silica, which exist in the earth's crust or in the atmosphere; in short, to those bodies in whose formation no living process had taken part, no seed had sprung to flower or fruit.

Our present-day definition of organic chemistry. — That was the sharp division that existed until Wöhler in 1828 made in his laboratory a typical organic compound which was built up wholly from inorganic matter without the aid of any life process whatsoever. By that one synthesis Wöhler tore down the wall which kept the two branches of chemistry apart, and proved that there was no essential difference between organic and inorganic chemistry. To-day we find thousands of substances which are formed by nature through life processes produced without the agency of life by the chemist in the laboratory, and built up from materials found in the earth's crust or in the atmosphere. Yet the old division has seemed so convenient a one for many reasons that it has persisted to some degree. To-day our new definition of organic chemistry is based on the fact that one element, carbon, appears in every natural compound produced by life forces, and in those laboratory compounds which resemble them. Our general definition can bear repeating once more. Organic chemistry is the chemistry of the carbon compounds. This definition is hardly close enough to fit all cases, as some carbon compounds, like carbon dioxide and the carbonates, belong especially to inorganic chemistry, and appear, as does carbonate of calcium (limestone or marble), in the earth's crust. These bodies act far more like inorganic substances. One chemist gives an excellent broader definition which includes such exceptions: "Organic chemistry is the chemistry of those compounds of carbon which contain in the molecule an atom of carbon, directly associated with either hydrogen, nitrogen, or another carbon atom." All such substances are grouped under the heading of organic chemistry.

The differing chemical characteristics of the two divisions of chemistry. — Probably the strongest reason for the broad separation into two classes is found in the differing chemical characteristics of the divisions. Why these differences occur we do not know, but we have proved by numerous examples that they do exist. Most of the inorganic compounds are good conductors in solution, and are broken down by the force of the electric current. Under these circumstances they separate in such a way as to make us believe that they are composed of positive and negative parts. Most organic compounds do not conduct electricity at all and are unaffected by its action. Organic compounds possess almost infinite powers of adding to themselves, adding like atom to like atom and like radical to like radical, thus forming whole series of closely related substances. Inorganic compounds possess very limited powers in this respect.

The five common elements found in organic compounds. —

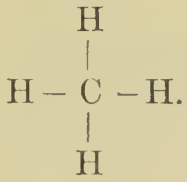
Beside carbon the vast majority of organic compounds contain but four other elements, hydrogen, oxygen, nitrogen, and sulphur. The compounds of inorganic chemistry, on the other hand, commonly range over three or four times as wide a field of choice and embrace almost every known element. The chemistry of organic compounds is chiefly non-metallic. The chemistry of inorganic compounds has more to do with the metals and their compounds than with any other class. The substances which are organic are chiefly light, low-boiling liquids and gases. Those which are inorganic are chiefly comparatively heavy solids, or heavy liquids. Typical inorganic substances may be found among those we have already studied. They are such bodies as sulphuric and hydrochloric acids, the metallic nitrates, sulphates, and carbonates, the hydroxides and ammonia. Typical organic substances will be found in the work of the next four chapters, which treat of alcohol and ether, chloroform, starch, sugar, and similar substances.

One point can hardly be insisted upon too forcibly. The difference between inorganic and organic substances depends chiefly on the characteristics of a single element — carbon.

Methane. — Suppose now that we pass from our general comparison of these groups to our consideration of some specific compounds. We naturally begin with methane, a compound of carbon and hydrogen, which is the first of the great series known as the hydrocarbons.

Occurrence of methane. Formula for methane =

CH_4.



$$\begin{array}{c} H \\ | \\ H-C-H. \\ | \\ H \end{array}$$

In the dark, gray days of later autumn, if you stand beside some forest pool burdened with fallen leaf and dying flower, you may see big bubbles of gas rising slowly to the surface, there to expand and break. If you touch a match to those bubbles you find that they spring into sudden flame. The gas within the watery film is methane, known for two centuries under the common name of marsh gas. Methane produces much of the flame which lights the heavens when a roaring natural gas well is ignited. Methane is the deadly fire damp of the mines, the gas whose explosion has cost so many miners' lives, and it is the gas which gives a large proportion of the heat energy furnished by the coal gas pouring through the meter in the cellar.

Decomposition. — The evolution of methane in a stagnant woody pool is typical of its general production in nature. It is formed when decomposition occurs with an insufficient supply of oxygen. The plant or leaf which dies in upper air forms carbon dioxide to a considerable extent as it perishes, because it can obtain sufficient oxygen from the stores of the atmosphere for complete oxidation. The dying leaves in the pool when they decompose form marsh gas, made up of carbon and four parts of hydrogen, because they are covered away from the oxygen of the air, and are therefore shielded from oxidation.

The value of carbon. — In the formation of methane, in the fact that in this compound one atom of carbon unites with four atoms of hydrogen, we find a condition which controls literally thousands of compounds in organic chemistry, the valence of carbon. This is a subject which has been discussed at greater length before, but slight reference to it here seems imperative. Generally speaking, by valence is meant the power of any element to combine with or replace the atom of hydrogen. Carbon in methane unites with four atoms of hydrogen, and is said to have a valence of four, or to be quadrivalent. One atom of chlorine unites with one atom of hydrogen in hydrochloric acid gas. Chlorine is said to have a valence of one or to be monovalent. One atom of sulphur unites with two atoms of hydrogen in hydrogen sulphide gas. Sulphur is bivalent. It has a valence of two. Iron in one of its chlorides replaces three hydrogen atoms. Here it is trivalent, it has a valence of three. Carbon is commonly quadrivalent as we have seen. That is, it combines regularly with four atoms of hydrogen or its equivalent. In carbon dioxide it combines with two atoms of oxygen. Two atoms of hydrogen combine with one of oxygen in forming water. Therefore two oxygen atoms must be equal to four hydrogen atoms, and carbon, which unites with two atoms of bivalent oxygen or four atoms of monovalent hydrogen, has a valence of four. We shall find that this quadrivalent property of carbon determines the structure of many a compound.

Preparation of methane in the laboratory. — Methane is made in the laboratory from our old acquaintance, sodium hydroxide, combined with another sodium compound which we have not met before, sodium acetate, the sodium salt of acetic acid. Marsh gas (methane) is produced when dry sodium acetate and sodium hydroxide are brought together in the presence of some substance like quicklime (calcium oxide), a body which will absorb any water formed. One atom of the two carbons in the sodium acetate and three of its hydrogen atoms unite with the one hydrogen atom of sodium hydroxide to form methane and leave sodium carbonate behind. That is:

Sodium acetate + sodium hydroxide = sodium carbonate + methane.

Formula for sodium acetate = $NaC_2H_3O_2$ or $(CH_3{\cdot}COO)Na$.

$$NaC_2H_3O_2 + NaOH = Na_2CO_3 + CH_4.$$

or

$$(CH_3 \cdot COO)Na + NaOH = Na_2CO_3 + CH_4.$$

Organic formulæ are very often written structurally with all the four bands of carbon shown by light lines, as in the formula of methane given three paragraphs back. When they are not written in that way, the metals or hydrogen in the compound are generally written last, reversing the order generally followed in inorganic chemistry. So sodium, acetate, the organic salt, is written $(CH_3 \cdot COO)Na$ where sodium sulphate, the inorganic salt, is written Na_2SO_4.

The derivatives of methane. — Despite the fact that methane readily burns in air with a hot flame, forming carbon dioxide with its carbon, and water with its hydrogen, it is, in general, a comparatively inert substance reacting with but few inorganic substances. But the number of organic compounds that are built up with methane as a foundation is legion. Some of them we shall consider here.

Wood alcohol. — Wood alcohol (methyl alcohol), known also as wood spirit, is probably the best known derivative of methane; it is, moreover, one of the simplest. When the charcoal burner chars his logs by destructive distillation, a series of liquids are formed which come from the complex substances that composed the wood. Of these methyl alcohol (wood alcohol) $(CH_3 \cdot OH)$ is one. In nature this compound is found in oil of wintergreen. Invariably poisonous, it has been the cause of numerous deaths. Yet its great power of solution makes it invaluable as a solvent for varnishes and dyes, for fats, oils, and resins. Boiling at 66 degrees C. it evaporates readily in the air, leaving behind the dissolved solids, which remain spread over the surface in an even smooth coating. This is the action which takes place in varnishing. The dry varnish dissolved in the alcohol is evenly distributed over the polished table top or carriage door. The solvent evaporates, dries, and the varnish coming out of solution remains as a solid coating.

The burning of wood alcohol. — Wood alcohol, like spirits of wine, our common alcohol, is colorless, lighter than water, and burns with a flame; it is, in consequence, a valuable fuel. When burned it breaks down into carbon dioxide and water, for it is made up of carbon, hydrogen, and oxygen. That brings us to our next point, the question of the chemical structure of this compound.

The radical methyl. — We cannot form compounds of methane by direct addition. When the carbon atom of methane has joined with its four hydrogen atoms the valence of carbon has been satisfied; it can combine with four hydrogen atoms and no more. In consequence we form derivatives of methane by removing hydrogen atoms from the compound and replacing them with the other atoms or radicals. The radical which is produced from methane by the removal of a single hydrogen atom is made up of one carbon atom and three hydrogen atoms. It is called methyl (methyl radical = CH_3). When we add a chlorine atom to this radical we have methyl chloride; an iodine atom joined to methyl gives us methyl iodide. The addition of a hydroxyl atom gives us methyl hydroxide. When we add a hydrogen atom to methyl we get methane.

Methyl hydroxide, wood alcohol. — Methyl alcohol like the other alcohols is a hydroxide, the hydroxide of methyl. The theoretical way of forming an alcohol is to replace one of the hydrogens of a primary substance like methane with a single hydroxyl radical. Methane contains one carbon atom and four hydrogen atoms. Alcohol or methyl alcohol is methyl hydroxide, made up of one carbon atom, three hydrogen atoms, and one hydroxyl radical.

Organic acids, bases, and salts. — In all the cases we have studied in inorganic chemistry, a hydroxide was a base and neutralized an acid. Though alcohols are unlike these bases (for example they do not color litmus), yet they have one of the essential properties that characterize the inorganic group. They form neutral bodies when they are treated with acids. The organic acids acting on the basic alcohols form neutral organic salts.

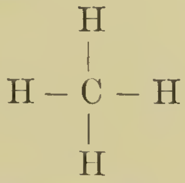
The aldehydes. — Next to methyl alcohol, the hydroxide of methyl, comes a group wholly unlike anything we have met before, the aldehydes. Strong as is the influence of the radicals (the groups of atoms acting as single atoms) in inorganic chemistry, the influence of the radical in organic chemistry often appears even more marked. As the alcohols are dominated by the hydroxyl radical, so are the aldehydes dominated by their characteristic radical, which is made up of one atom each of carbon, hydrogen, and oxygen. All the aldehydes contain this radical. The aldehyde radical = (CHO).

Formaldehyde. — Formaldehyde (methyl aldehyde) (H·CHO) or formic aldehyde is the best known of all the members of this group. It is a deadly foe to the bacterial life of the air and of the soil, and so is widely used as a germicide and disinfectant. It is also a valuable medical agent, for it destroys the micro-organisms of disease, and the ease with which it can be handled, transported or sprayed makes it of great value. Formalin, the common preservative of the biological laboratory, is a water solution of this compound.

The germicidal uses of formaldehyde. — The ease with which formaldehyde goes into solution, and its powerful action as a germicide, have brought about some evil as well as some good effects. The souring of milk or the decay of meat is accomplished by the agency of tiny microscopic plants and animals. These germs or micro-organisms, as they grow and multiply, produce chemical changes tending in the end to completely change the substance in which they lodge. Most germs are killed by formaldehyde, so the dealer whose milk is liable to be soured, or the meat man whose meat has been too long on hand, has in the past sometimes yielded to the temptation to keep his milk or his meat by the addition of this liquid. Formaldehyde in any quantity is believed to be injurious when taken into the body through the alimentary canal, excellent as a dilute solution of it is when used externally. Hence the danger of formaldehyde-preserved foods. In many of the states the use of formaldehyde as a preservative is forbidden by law, and one of the regular duties of the boards of health is to make analyses which shall insure a clean supply of food unchanged by preservatives.

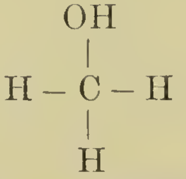
Preparation of formaldehyde. — Formaldehyde may be produced by the gradual oxidation of methyl alcohol, as when alcoholic vapor mixed with air is passed over heated platinum. This action is in many respects similar to the oxidation of sulphur dioxide to sulphur trioxide by the contact process. In the formation of formaldehyde from alcohol, however, two products are formed. Water is produced beside the oxide. Formalin is tested by the silver mirror formed when the compound is added to an ammoniacal solution of silver hydroxide.

The oxidation of methane. — The gradual production of methane derivatives may be illustrated by the following diagram:



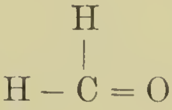
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Methane CH_4



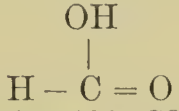
$$\begin{array}{ccccc} & & OH & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Methyl alcohol $CH_3 \cdot OH$



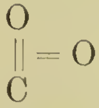
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & = & O \end{array}$$

Formaldehyde $H \cdot CHO$



$$\begin{array}{ccccc} & & OH & & \\ & & | & & \\ H & - & C & = & O \end{array}$$

Formic acid $H \cdot COOH$



$$\begin{array}{cl} O & \\ \| & = O \\ C & \end{array}$$

Carbon dioxide CO_2

Each of the derivatives shown above is produced by the addition of a single oxygen atom. When one is added a hydroxide is formed. When two oxygen atoms are added and two hydroxyls would be expected to form, water drops off and an aldehyde is produced. A single atom of carbon will not hold two hydroxyl radicals. Water is formed when two hydroxyl radicals thus occur. When another oxygen atom is added to the aldehyde, the radical carboxyl (COOH) forms. When still another atom of oxygen is added and two hydroxyls would naturally result, water drops off and a molecule of carbon dioxide is left as the final product of oxidation.

Formic acid. — Methyl alcohol, the hydroxide of methyl, is theoretically produced by adding one atom of the element oxygen to methane. The addition of this atom produces the radical hydroxyl. Formaldehyde, the aldehyde derivative of methane, is theoretically produced by the addition of a second oxygen atom. If we add a third oxygen atom we find that a third compound is produced by successive oxidations. This substance is the first of the wide-spreading group of organic acids, formic acid. As the characteristic radical of the alcohols is hydroxyl, made up of one atom of hydrogen and one of oxygen, and as that of the aldehydes is made up of one atom of carbon, one of oxygen, and one of hydrogen, so here we have a radical characteristic of the acids and named carboxyl. The carboxyl radical which belongs especially to the organic acids is made up of one carbon atom, two oxygen atoms and one hydrogen atom. The sting of the nettle and the bee, the bite of the red ant, all irritate and poison because of the presence of formic acid. In bygone years formic acid was regularly prepared by macerating the nettle of the pasture in large vessels and afterwards distilling the macerated mass. From this distillation arose a stinging, irritating, gaseous acid which, when condensed, fell into the receiver as a moderately heavy, colorless liquid with a penetrating odor. Dropped on the skin formic acid causes painful wounds and raises blisters. When cooled it solidifies to a crystalline body. It boils at 99 degrees C., just one degree below the boiling-point of water. In the laboratory this substance is produced from oxalic acid. Oxalic acid, when heated, gives formic acid and carbon dioxide. A whole series of salts called the formates are produced from formic acid.

Methyl ether. — The oxide of methyl is an ether. Like the well-known sulphuric ether, used as an anæsthetic, this compound is prepared by the action of sulphuric acid on an alcohol. Instead of using spirits of wine (the alcohol which gives our ordinary ether), wood alcohol (methyl alcohol) is used. The action of the sulphuric acid found here is one which is commonly used in organic chemistry, and depends on the dehydrating or hygroscopic power of the acid, that is, on the removal of sufficient hydrogen and oxygen from a compound to form water. When two parts of methyl alcohol are acted upon by sulphuric acid, water drops off and methyl ether (the oxide of methyl) [$(CH_3)_2O$] is left behind.

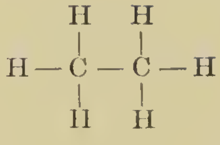
Chloroform and Iodoform. — Chloroform (tri-chloromethane = $CHCl_3$) and iodoform (tri-iodomethane = CHI_3) are chloride and iodide derivatives of methane. Chloroform will be mentioned later under our brief study of the anæsthetics, but as these two substances are directly derived from methane they should be mentioned here. They are typical organic halogen compounds. One point should be mentioned before the chapter closes. In writing the formulæ of organic compounds, a period often divides two radicals. So the formula of acetic acid written as $CH_3 \cdot COOH$ (or $CH_3 \cdot COOH$) indicates the presence in the compound of carboxyl and the methyl radical.

The hydrocarbons as a whole. — The study of methane and its derivatives gives us the key which unlocks the first great division of organic chemistry, the division known as the hydrocarbons. A cipher letter, a cryptogram in which stand letters and numbers mingled in disorder, may blind and utterly confuse the chance reader. So the compounds of organic chemistry mounting in number to the thousands seem a bewildering tangle to the student at his first approach. But a cipher letter or a cryptogram has a key. To the man who knows that key the letter or the puzzle falls at once into coherency and sense. In the same way organic chemistry has a series of master keys. Hundreds of alcohols, ethers, aldehydes, or acids can be produced from compounds which resemble methane. These are derivatives of hydrocarbons, formed, like methane, of carbon and hydrogen. If we know how the compounds of methane *are* formed, we know how the compounds of every substance like methane *will* be formed. Every alcohol of the long list is, like methyl alcohol, a hydroxide. Every alcohol contains the aldehyde radical. Every acid contains the carboxyl radical. Every ether is an oxide. Each member of each group is marvelously like every other member. If we know the links of the methane chain we know the more essential facts of the links that make up every other chain. We hold the opening key.

CHAPTER XXVI

ETHANE

Formula for ethane = C_2H_6 ·



$$H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H$$

Ethane and its derivatives. — The never dying flame which rose upon the ancient altars of the fire worshipers finds its nearest natural parallel in the unceasing vigor of the roaring fountains of fire that rise from the natural gas wells of the petroleum lands. Among the mingled gases which produce that flame is found ethane, the second member of the hydrocarbons. Methane and other compounds are found there as well, but whatever importance these other constituents may possess in themselves, none have given more important compounds to the world than ethane. This colorless, tasteless, insoluble gas is lighter than air, and burns with a slightly luminous flame. From it are derived many substances without which our modern life would be indeed impoverished: alcohol, the alcohol of medicine, of stimulants and of fuel; ether, the merciful anæsthetic that deadens the pain of life-giving operations; vinegar, most used of all the food acids, whose tartness comes from the acid derivative of ethane (acetic acid); esters, whose presence gives the delicious odor and taste of many a fruit and flower; acetates, which dye rustling silk and smooth fresh cotton.

The chemical constitution of ethane, of propane, and the other hydrocarbons of this series. — Methane is made up from one atom of carbon joined with four atoms of hydrogen. Ethane is formed by the addition to methane of another carbon atom and two hydrogen atoms, so that this compound holds altogether two carbon and six hydrogen atoms. Every like member of this series, known as the hydrocarbons, is formed in a fashion similar to the production of ethane from methane, that is, by the addition of a carbon and two hydrogen atoms to the carbon and hydrogen of the member of the group which just precedes it. Propane, next above ethane in the series, is made by adding another carbon and two more hydrogen atoms to ethane, and therefore contains three carbon and eight hydrogen atoms. The list of compounds so formed might theoretically be greatly extended. A German mathematician once figured out the space which the formulas of possible compounds such as these would occupy if written in massive volumes. He decided that the Houses of Parliament at Westminster, if packed from top to bottom with folios, would not hold them all. Every compound of this series of the straight chain up to one which contains sixty carbon atoms and one hundred and twenty-two hydrogen atoms has been made in the laboratory by adding again and again a single carbon and two hydrogen atoms in a straight line. We need trouble ourselves but little concerning the higher forms, however. All of them act in much the same way as methane and ethane, and many are of but little practical importance.

Ethyl alcohol. Our ordinary alcohol. — As wood alcohol (methyl alcohol) is the hydroxide of methyl, so spirits of wine (ethyl alcohol) is the hydroxide of ethyl, an organic radical formed from ethane by the removal of a hydrogen atom, and containing, therefore, two carbon and five hydrogen atoms. This is the alcohol of common use. As we shall see in the next chapter, alcohol is produced by fermentation, by a phenomenon which happened alike in the dawn of history and yesterday, a process as much a part of nature's manufactories as is the growth of bark on a tree. Like methyl alcohol from methane, so ethyl alcohol can be obtained from the oxidation of the compound ethane from which it comes. Here, however, as in many another place, the bounteous wealth of nature's factories gives by far the easiest method of production. The sugar-cane and the sugar-beet can furnish sugar far better than any synthetic laboratory process. Alcohol has been produced from fermented fruit during the recorded existence of man. The old process controlled by modern methods governs the manufacture of alcohol to-day.

Fractional distillation. — Fermentative processes will be considered in the next chapter. It is sufficient here to say that alcohol is produced by fermentation, and that it comes to us from these processes impure. To purify it we must divide the mixture and separate the component parts. Fractional distillation accomplishes this. The purification of a liquid by this method depends on the fact that most liquids boil and turn into gases at different temperatures. Pure alcohol, for example, boils at 78 degrees C., water boils at 100 degrees C. Suppose we heat a mixture of water and alcohol to 78 degrees C. There the alcohol will begin to come over as a gas. Raise the temperature a little higher and a gaseous mixture of water and alcohol comes off. Raise the solution to 100 degrees C. and the whole distills over. As the temperature climbs higher and higher from 78 degrees C. to 100 degrees C., more and more water, less and less alcohol, will come over. All the vapors of every kind, which come from the flask holding the mixture, can be changed to liquid by passing them through a cooled tube exactly as water can be condensed. The condensed liquids can be separated by collecting them in different vessels, the strong in one and the weak in another. The distillate, as it is called, which comes over around 80 degrees will be rich in alcohol. That which comes over at the higher temperatures will be poor in alcohol. None of the distillates will be pure alcohol. If we divide the distillation into a half dozen portions, taking for the first the liquid we collect while the thermometer rises from 78 degrees to 82 degrees, and for each of the others the amount which comes off while the liquid rises 4 degrees, we shall have at the end half a dozen flasks. The first will be strong alcohol. The last will be almost pure water. Now if we repeat this process with each of the collected portions we shall get finally a division by which practically all the alcohol will be collected in one flask and all the water in another, for each time we distill we get more alcohol over at the lowest temperature until, in the last, we get a liquid practically all of which boils at that temperature. We can never wholly separate alcohol from water in this way; to do that requires more extended treatment, but we can greatly purify it. Many organic liquids can be completely separated and purified by this process. When we obtain a liquid which wholly boils away at a single temperature, that is, which has a constant boiling-point, we may feel sure that we have a pure substance. This statement is equally true of those substances which have a constant melting-point. Pure substances are the only ones which show these peculiar properties.

The properties of alcohol. — Alcohol may be separated from fermented mixtures of fruit juices by the process of fractional distillation just explained. When obtained it is a colorless liquid about four fifths as heavy as water, which, though it boils at the low temperature of 78 degrees C., does not solidify until the cold of -130 degrees C. is reached. Two properties, namely, that alcohol remains liquid under the influence of great cold, and that liquid alcohol expands and contracts with regularity as it meets warmth or chill, makes spirit thermometers containing this compound of much value in northern regions. In the steel-blue and stone-gray of Arctic ice fields, a mercury thermometer would contain a piece of solid metallic mercury which would not vary in volume with change of temperature. Alcohol will not freeze when exposed to the lowest temperature of the lands that border on the poles, so the red spirit thermometer gives true readings in the frozen north. An interesting modern use of the fact that alcohol freezes only with difficulty is found in the fact that automobilists place alcohol in their cooling water in winter. The presence of the alcohol keeps the water from freezing.

The burning of alcohol. — Alcohol burns with a pale, nonluminous flame, and, like methyl alcohol, forms an explosive mixture with air or oxygen. When burned, it follows the general rule of the alcohols and unites with oxygen to form carbon dioxide and water, its carbon combining with oxygen to form the first and its hydrogen combining with the same element to form the second.

The solvent powers of alcohol. — Like wood alcohol, spirits of wine is an excellent solvent. Both of these alcohols have neutral reactions, but both form salts with acids as if they were weak bases.

A test for alcohol. — The presence of alcohol can best be detected by the formation of that valuable antiseptic used for surgical dressings, iodoform. If we place a small quantity of iodine in a solution where alcohol is present, and, after heating gently, add potassium hydroxide a drop at a time, we shall find that a yellow precipitate forms slowly. The six-sided crystals of yellow iodoform produced by the action may be recognized by their peculiar clean surgical odor and by the appearance of the crystals.

Sulphuric ether. — Sulphuric ether, best known of all the anæsthetics, has stood for more than sixty years as a guard against pain. Before the days when unconsciousness was procurable during surgical operations many lives were unnecessarily lost. The pain and shock of operations were too great to be endured, and the involuntary movement of the sufferer made delicate manipulation impossible. With anæsthetics came the power of producing anæsthesia at will. Anæsthesia may briefly be defined as the production of a state of insensibility to external impressions.

Anæsthetics and the discovery of sulphuric ether. — The conception that it was possible to produce anæsthesia artificially came first from the teeming imagination of Sir Humphrey Davy, who, in the early part of the century, proposed the use of nitrous oxide. He based his suggestion upon some accidental discoveries of the anæsthetic power of laughing gas (nitrous oxide). Scarcely anything was done, however, until 1844, when Dr. Wells, of Hartford, Conn., used nitrous oxide in an experimental trial. This gas, though available for small operations, was not sufficiently lasting for greater ones and did not secure sufficiently complete results. It was not until 1846 that the new era of surgery opened. In that year Dr. Morton of Boston used sulphuric ether successfully for the first time, and the discovery of its anæsthetic value was soon followed by the discovery of the possibilities of chloroform. Through those discoveries came the blessing of unconsciousness during surgical operations. It was proven that anæsthetics could be used practically without danger when skilfully applied, a proof which was in reality quite as essential as the primary one of anæsthesia. The lowering of the death rate by modern surgery is largely based upon the work of the pioneers in anæsthetics.

SIR HUMPHREY DAVY. (1778–1829)

A great English chemist, theorist and discoverer of several elements.

The ethyl radical. — Sulphuric or ethyl ether is an oxide of the ethyl radical (C_2H_5). This radical, as we have already seen, contains two carbon and five hydrogen atoms. It is monovalent, that is, it can unite with one hydrogen atom or its equivalent. When ethyl unites with one hydrogen we have the primary compound ethane made up of two carbon atoms and six hydrogen atoms (C_2H_6). When it unites with hydroxyl instead of hydrogen we have ethyl hydroxide or ethyl alcohol ($C_2H_5{\cdot}OH$). When we form the oxide we form it by the use of two ethyl radicals joined, as oxygen is bivalent $(C_2H_5)_2O$. So ether is made up of two ethyl radicals joined with oxygen, and all the ethers of the series follow the general rule observed in this instance.

The preparation of sulphuric ether. — Ethyl ether is made from ethyl alcohol by the action of sulphuric acid, whence it derives its name sulphuric ether. Its preparation depends largely upon the power of sulphuric acid to take water from any substance containing it. When hydrogen and oxygen in the proportion to form water are taken from two molecules of ethyl hydroxide (ethyl alcohol) ($C_2H_5 \cdot OH$), an oxide (ethyl ether) $(C_2H_5)_2O$ is left behind.

The properties of ether. — Ethyl ether is a very volatile and highly inflammable substance. Placed on the hand it evaporates so rapidly that the skin feels cold from the heat lost in carrying on the evaporation. The inflammability of this substance is so great that all experiments in which it is used should be carried on at least ten feet from any flame. It should only be heated on a steam bath. Ether is a clear liquid, about three fourths as heavy as water, and is a great solvent, putting waxes, greases, oils, etc., into solution. This special quality of ether leads to its employment in fluids used in cleansing clothes. The inflammability of the liquid makes such cleansers dangerous when they are exposed to a flame. Care should be taken never to have a flame near a cleansing solution which smells of ether or of similar substances. Many serious fires have come from the neglect of this simple precaution.

The esters or ethereal salts. — The esters or ethereal salts which are formed from ethyl alcohol should receive brief mention here. These compounds are found in many of the fruits and flowers, giving them peculiar flavors or fragrances. Many of them can be prepared in the laboratory. A type of these bodies is ethyl acetate, a fragrant compound smelling like cider, which is produced by the neutralization of ethyl alcohol by acetic acid. In general these substances are prepared by the neutralizing action of an alcohol upon an organic or an inorganic acid.

The destructive distillation of wood. Pyroligneous acid. — When wood is destructively distilled, that is, when it is heated in closed retorts away from possible contact with the oxygen of the air, four products are commonly obtained, gases, thin liquids, wood tar, and charcoal. The gases are such bodies as hydrogen, methane, ethane, and the like. These may be burned as fuel. The solid matter is largely charcoal which we have already discussed. Wood tar is used for the production of light oil used as a substitute for oil of turpentine, and of heavy oils for use with axle greases and lampblack. Of the thin liquids the most important is one known as "pyroligneous acid." This liquid distills over from heated wood between the temperatures of 160 degrees and 275 degrees C., and contains three important organic compounds, wood alcohol, acetone, and acetic acid. The first, wood alcohol (methyl alcohol), has already been discussed. The second, acetone, $(CH_3)_2CO$, is a solvent which is typical of the group known as the ketones. This class of substances is characterized by the radical carbonyl (CO), a radical composed of one atom of carbon and one of oxygen. The third compound, acetic acid ($CH_3 \cdot COOH$), is one of the great organic acids. In some respects it is the greatest. The relation of its structure to that of the other organic compounds of its series will be discussed in the chapter on Fermentation, where the connection of this substance with vinegar is considered. It is sufficient here to speak of the common way of obtaining the acid and of the use of its compounds, the acetates.

The preparation of acetic acid. — When pyroligneous acid, which is a mixture of several liquids, is passed through lime the acetic acid present combines with lime to form calcium acetate. This is a typical case of the neutralization of an acid by a base. The reaction which takes place is as follows:

Calcium hydroxide + hydrogen acetate = calcium acetate + water.

$$Ca(OH)_2 + 2CH_3 \cdot COOH = (CH_3 \cdot COO)_2\,Ca + 2H_2O.$$

The calcium acetate when formed is in solution. This solution is filtered and evaporated to dryness. It is then used as a basis for preparing the commercial acetic acid. For this purpose the calcium acetate is distilled with concentrated hydrochloric acid in great copper stills. Acetic acid and calcium chloride are formed.

Calcium acetate + hydrogen chloride = calcium chloride + hydrogen acetate (acetic acid).

$$(CH_3{\cdot}COO)_2\ Ca + 2HCl = 2CH_3{\cdot}COOH + CaCl_2.$$

The acid obtained contains about fifty per cent of anhydrous acetic acid. This can be concentrated to a solid known as glacial acetic acid. Acetic acid is used in making acetates, in preparing white lead, in medicine, and in the manufacture of dyes and explosives. It is, as we shall see, the acid principle of common vinegar.

The acetates and their uses. — The acetates are metallic salts of acetic acid. Several of them are produced by acting on other types of metallic salts with lead acetate, a compound prepared by dissolving an oxide of lead in acetic acid. Here is a typical reaction for preparing aluminum acetate, a salt much used in dyeing:

Aluminum sulphate + lead acetate = aluminum acetate + lead sulphate.

Many of these salts are used in dyeing. The acetates of aluminum, of chromium, of iron, of sodium, and of lead are all employed for this purpose, while a number of other acetates serve important uses in the preparation or separation of various chemicals used in the arts.

CHAPTER XXVII

FERMENTATION

The garden of the air. Bacteria, yeasts, and molds. — Our summer garden fades. Bright flower petal and glossy leaf alike shrivel and die when winter chills the soil. But all about us, summer and winter, in cold and heat, in moisture and in drought, flourishes another garden — the garden of the air, whose plants are the tiny forms known as the bacteria, yeasts and molds. The room in which you sit holds thousands of them floating in its atmosphere, lying in the dust of the window-sill or the corner, growing swiftly in any warm, moist substance, such as a bit of bread that may have been overlooked by the housekeeper's eye.

Occurrence of the molds. — Invisible as are most of these individual plants when separate, they may be recognized when brought together in a mass. The molds are the largest of the three divisions, and every one has seen their white threads growing on bread, apple, or banana, or has noticed their masses spreading over cheese, white at first, but turning bluish green when the microscopic fruit appears. A hot, humid summer which provides two requisites most favorable to plant life, warmth and moisture, finds this mold growing on leather boots put aside for but a few days.

Yeasts. — The yeasts are known to every man, woman, and child in the form of compressed yeast cake, sold for bread-raising. A yeast cake is simply a mass of yeast cells which remain for some time living but quiescent when kept dry and cool (hence the necessity for keeping yeast cakes in the ice-chest), but which grow readily when placed in the warm, moist dough, producing by their growth the raising of the bread. As in the case of the molds the yeasts follow the general rule of most plant life. They grow rapidly in moisture and warmth. They live quiescent in dryness and cold, and die when they are exposed to such conditions for too long a time. Of the three forms we are considering, yeasts are the most helpful to man.

Bacteria and their characteristic properties. — The bacteria are least known of all the forms except where they appear as "mother of vinegar" (the plant life which appears in the vinegar cruet in masses containing millions upon millions of individuals). Those who have peered through the microscope at bacteria are generally the only ones who have seen different types of these infinitesimal plants, many of whose forms could rest in numbers on the extreme point of the finest cambric needle. Fifty thousand small bacteria could lie in a straight line across the inch space on a ruler. The largest that are found do not exceed a ten-thousandth of an inch. Despite their small size, however, the work that they do is a mighty one both for good and evil. These plants are the scavengers of the air, falling on waste organic matter, decomposing it and leaving harmless inorganic wastes behind. Their action in this respect was noted in the study of sewage disposal, where it has already been stated that some of the best modern sewage plants are piles of rocks whereon rest bacterial films. Sewage poured over these piles serves as material to foster bacterial plant life, and after giving up its dangerous contents as it passes through, it flows away harmless.

Were it not for the decay produced by bacterial growth, life might have long since ceased upon this planet. The reason for this strange statement follows: To produce life-giving foods from the soil, the woody fiber of trees, which, in its ordinary form, is useless for plant food, must be brought by decay into a condition where it can be made fit nourishment for plants. It is micro-organic growth which prepares the food of the plant from the fiber of the tree. Were it not thus prepared the stock of available organic carbon would before this have been largely turned to harsh unyielding wood incapable of giving nourishment to man. Decay, the rotting of the apple or the tree, is due to the softening and destruction of organic matter by means of the growth of micro-organisms. It is a part of the wise provision of nature that the waste products of the earth should be removed and kept from cumbering the soil. Decay in the end means the removal of dead wastes, for final decay brings the substance back to those simple, inorganic, harmless compounds from which it sprung — to carbon dioxide, to water and the like.

Processes dependent upon bacterial growth. — Many of the processes which the old-fashioned farmer believed occurred by chance are recognized by the scientific farmer to be really due to the growth of bacterial plant life under conditions of heat and moisture favorable to their increase and continued existence. The modern agricultural student knows that "mother of vinegar," which brings vinegar from his cider, is a type of this group, and that the flavors of his butter and cheese are produced by the presence of special kinds of bacteria. He recognizes that though it is true that many bacteria cause decay and that a few cause serious disease, yet there is no more reason for hesitating to use helpful bacterial plants in producing foodstuffs than there would be in refusing to mix tomatoes and lettuce together to form a salad. In the former case we have small plants grown in healthful surroundings. In the latter case we have large plants grown in healthful surroundings.

General characteristics of the micro-organisms. — The molds, yeasts, and bacteria are all grouped under the general headings of micro-organisms or microbes. These bodies differ from the ordinary plants of the soil chiefly in the way in which they procure their food. In this they follow the habit of their general class, the fungi. The green plants, such as the grasses and cereals, vegetables and flowers, take in the carbon dioxide of the air, and use its carbon by means of the green coloring matter, chlorophyll, which is in their component cells. The colorless plants, the fungi, do not possess this chemical reagent, cannot obtain their food from the air, and must obtain it from the soil where they grow. The toadstools and mushrooms, like other types of fungi, obtain their nourishment from the organic matter on which they rest. The mold which forms on cheese or jelly, the yeast which grows and thrives in fermenting molasses, or the bacteria which produces vinegar, all draw their sustenance from the spot where they lodge, and thrive most when they meet three favorable conditions, nourishment, warmth, and moisture. All three of these would be afforded a mold, for example, on a piece of fresh cheese at the temperature of an ordinary house.

Micro-organisms in their relations to the foods of man. — In general micro-organisms meet conditions favorable to their growth in the foods of man. Acting upon the foods they change the chemical constituency of these compounds to a marked degree, and while they consume them they commonly make them unpalatable. In consequence, a large part of the science of the preservation of food consists of sheltering it from the attacks of harmful organisms. The cook, knowing that moderate heat and moisture injure food, keeps foodstuffs in a dry ice-chest, covers her jellies with a paraffine cover to shut out the air, heats her preserves to boiling to kill the germ life that may exist within, and adds heavy quantities of sugar to her jellies (for bacteria do not grow readily in pure sugar or in heavily sugared solutions). Curing factories salt ham and bacon, pork and corned beef, for like reason of preservation. Mackerel is salted for transportation. Fresh butter is made salt for keeping. All this is done to get rid of growing organisms. Germ life cannot live in salt or where salt is largely present. It should, however, be remembered that necessary as salt is, and much as salt foods enter into our diet, that the body demands fresh foods for its chief sustenance. Those who have tried to subsist on salt food exclusively have met with baneful results.

Fermentation. — In the change of sweet cider to hard cider, and the subsequent change of hard cider to vinegar, we have an example of two typical processes carried on by yeasts and bacteria. We shall consider the first change here which takes place through fermentation carried on by means of yeasts, the fermentative workers of the world.

The reactions involved in fermentation processes are, in general, too complex for statement here. One reaction may be given. When glucose, one of the sugars, is fermented it breaks down into alcohol and carbon dioxide as shown here:

Glucose = alcohol + carbon dioxide.

$$C_6H_{12}O_6 = 2C_2H_5{\cdot}OH + 2CO_2.$$

Ordinary cane-sugar or sucrose has the formula:

$$C_{12}H_{22}O_{11}.$$

Wild yeasts and the change from sweet cider to hard cider. Fermented liquors. — Accustomed as we are to the domesticated yeast cake, it is sometimes hard for us to realize that there are always wild yeasts floating in the air as well as molds and bacteria. The change from sweet to hard cider is brought about by these wild yeasts. When we bruise and crush apples in a cider press we are mingling with the juice thousands of micro-organisms which have lodged on the skins. The yeasts which are among these organisms find the juice of ripe apples a particularly suitable place for growth, and increase in it rapidly. As they increase the peculiar process known as fermentation commences, and goes on as the yeast plants grow. Generally speaking, this process is carried on in the following way. All fruit juices contain sugar, and it is on the action of the yeasts on the sugars that the essential characteristics of the process depend. When the yeast attacks the sugar, which is made up of carbon, hydrogen, and oxygen, it breaks down into two simple parts, ethyl alcohol (spirits of wine) and carbon dioxide. The gas rises in bubbles through the liquid and the alcohol remains. Just that happens in the hardening of cider. When the wild yeasts that rested on the apple skin or dropped into the cider from the air commence to grow and take in food they begin to act on the sugar of the apple juice. Soon bubbles of gas, of carbon dioxide, begin to rise and continue to be evolved throughout the fermentation. As hard cider is produced little by little the percentage of alcohol grows greater. Alcohol is the second product which always appears in fermentative changes. In the change of fruit sugars to alcohol and carbon dioxide under the influence of growing yeast lies the principle of fermentation. In fermenting wines the wine grower depends largely on the yeast which lies on the skin of the grape. In breweries and distilleries, such cereals as rye, corn, barley, and the like, which have generally been previously malted, a process which turns a considerable part of their starch to sugar, are mixed with yeasts and fermentation is allowed to proceed in huge vats. Such action produces the lighter fermented drinks like beer and ale. Distillation from these fermented liquids produces the heavy distilled liquors like brandy and whisky wherein alcohol is concentrated and used in high per cents.

The individual yeast plant and the way in which it grows. — Much evil has been brought upon the world by the spread of alcoholic liquors, and many fearful problems arising from their use are still unsolved. It is pleasant, therefore, to turn to the other great use of yeasts and tell of the part they play in making the brown, crisp, sweet-smelling loaves of bread that came from the oven last baking day. Yeast plants are never seen as individuals with the naked eye. A cubic inch of compressed yeast can easily contain over twenty billion individuals. Their action in bread is due to great quantities of these plants all growing at the same time. A single yeast plant under the microscope commonly looks something like a tiny transparent egg, inside of which can be seen smaller bodies to some degree resembling an egg's yolk. These inner bodies are of two kinds, one called a vacuole, which is like a drop of oily liquid, a substance which has little apparent connection with the life of the plant, and a second body, the nucleus of the plant, which is one of the two necessary parts of all living cells except the lowest types. The other necessary part is the surrounding mass, the cytoplasm. Yeast grows by budding, that is, it grows by buds springing from its side, as buds appear on the boughs of a tree in spring. When we examine these buds we find that the nucleus of the original cell has divided between the old cell and the new so that both the bud and the plant possess nuclei. The yeast cake is made up of yeast cells, and grows in the dough of the bread raiser by budding. The necessity for freshness in a yeast cake appears in the fact that bread raising comes from living, growing plants, and the longer the cake is kept without nourishment the more individuals will die, just as geraniums taken from the ground would die without soil or water. Beside such growth by budding, yeast plants grow by forming spores, those tiny germs of the flowerless plants which carry on their life as do the seeds of the flowering races. The wild yeasts of the air come chiefly from spores.

The raising of bread by means of yeasts. — The raising of bread is produced by the growth of yeast plants by budding. Suppose we carry that action a little farther. We have already seen that yeast, causing the fermentation of a sugar solution, produces alcohol and carbon dioxide. Both those principles play an important part in bread-making. When the dough is first mixed there is some sugar in the flour, but not a sufficient quantity. To obtain enough it is necessary to change some of the starch of the flour into sugar, an operation which is carried on by means of a second substance present in the flour known as diastase. The combination of the sugar originally present plus the sugar newly formed from the starch of the flour produces a sufficient quantity to sustain the life of the yeast and encourage its growth. After the yeast has been mixed with moist dough by kneading it has two of the conditions most favorable for growth, nutriment and moisture. When the dough is put in a warm place to rise, the third condition, warmth, is attained. Common yeasts swiftly begin to bud in the dough, forming new plants which require nourishment. As the tiny organisms feed they cause the sugar to ferment, producing alcohol and carbon dioxide gas. The carbon dioxide which is produced cannot escape from the sticky mass, so it forms in bubbles (the pores of porous bread) inside the dough, forcing it to rise. When the dough has been again kneaded and risen it is placed in the hot oven, where the heat drives off the small amount of alcohol which has been formed, expands the gas confined in the pores of the dough, making larger holes, and, before the carbon dioxide gas escapes, the heat hardens the bread about the holes. As a result the bread is left porous and light. "Raised bread" (porous bread) is far better adapted to the digestive juices than unraised or unleavened bread.

The flavor imparted by yeasts. — According to some authorities the flavor imparted to bread by yeast plays no small part in making bread palatable. This flavor is lost when bread is raised by the soda process described in the chapter on the carbonates, or when (as in the so-called aerated bread) the pores are produced by mixing air with the dough in a machine. Yeast, the basis of the leaven of the Bible, caught from the garden of the air and turned to the making of leavened bread by the Israelites, remains the leavening power of most of the bread baked to-day as it has been for centuries past.

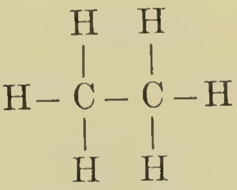
The change of hard cider to vinegar. The growth of bacteria. — The changes which cider undergoes in passing from sweet cider to hard cider are due to the fermentative action of the yeast. The change from hard cider into cider vinegar is caused by the action of the bacteria. We have already spoken of the minuteness of this third group of plants that flourish in the garden of the air. Their structure deserves a moment's notice. Far more abundant than the yeasts, the bacteria throng in the soil, abound in rich, dark earth, live in water, and float in myriads in the air. They increase with a rapidity almost beyond the limits of the imagination, since, if nothing were opposed to their growth, certain types of individual bacteria could produce from fifteen to twenty million of off-spring in a single day of twenty-four hours. Fortunately, the death rate of these germs is enormous, so that cold, lack of food, and similar obstacles keep the bacteria down to comparatively moderate quantities. Yeasts grow by budding; bacteria grow by simple lengthening and division. It is as if a single willow-tree stalk, six inches long, grew to a foot, then automatically broke in two and became two six-inch willow trees each of which went on lengthening to a foot and dividing into two like the first. Whether the bacteria be spherical, rod-like, or spiral in shape they all grow by lengthening and division.

Making vinegar. — Hard cider contains alcohol, and one type of bacteria, the "mother of vinegar," as it is called, thrives best in weak alcoholic solutions which contain some other organic matter for nourishment, such as apple-juice or apple-parings. Some organic matter of that nature must be present for vinegar making, since the bacteria will not live in pure diluted alcohol. In one process for making cider vinegar the cider is poured into casks filled with beechwood shavings and covered with colonies of "mother of vinegar." As the cider flows over the shavings the bacteria act upon it and change the alcohol to an acid (acetic acid), that acid of ethane which corresponds in structure to formic acid. This compound may be formed from beer, weak wines, or other impure alcoholic solutions as well as from cider, but when so formed it lacks some of the peculiar flavors which come from the apple juice.

Uses of acetic acid. — A large amount of acetic acid goes as vinegar into those pickles which housewives put up in the fall, or into those advertised in street car or on railroad line. Perhaps the commonest of the organic acids, as we have seen acetic acid is manufactured chiefly by the dry distillation of wood and is sold in a pure state either in dilute water solution or as a concentrated solid known as glacial acetic acid. It is a widely used solvent which often replaces alcohol.

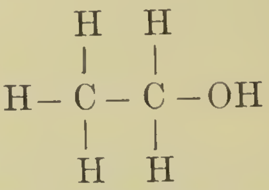
The oxidation of methane. — Acetic acid is another type of the organic acid of the hydrocarbons, one of the great series made from primary compounds like methane and ethane by successive oxidations. Oxidize methane and you have methyl alcohol (wood alcohol) containing one oxygen atom. Oxidize that and water drops off (no single hydrocarbon radical can hold two parts of hydroxyl) and formic aldehyde (formaldehyde) is left. When one more oxygen atom is added we have formic acid, while yet another atom causes water to drop off once more and the end reaction is reached in carbon dioxide. That gas is the common end of organic oxidation.

The oxidation of ethane. — By similar processes, that is, by the addition of single atoms of the great combining gas oxygen, we can change ethane to ethyl alcohol (spirits of wine), to ethyl aldehyde (acetaldehyde), to acetic acid, and lastly to carbon dioxide. For that matter we can provide an alcohol, an aldehyde, and an acid for any of the long list of the hydrocarbons by oxidations quite like these. The derivatives of ethane produced by oxidation follow closely the derivatives of methane as mentioned in the last chapter. The following formulæ may aid in making this clear:



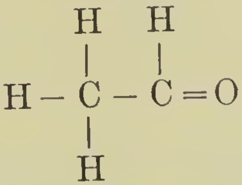
$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

Ethane C_2H_6



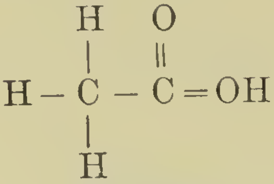
$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & OH \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

Ethyl alcohol $C_2H_5 \cdot OH$



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & = & O \\ & & | & & & & \\ & & H & & & & \end{array}$$

Ethyl or acetic aldehyde
$CH_3 \cdot CHO$



$$\begin{array}{ccccccc} & & H & & O & & \\ & & | & & \| & & \\ H & - & C & - & C & = & OH \\ & & | & & & & \\ & & H & & & & \end{array}$$

Acetic acid
$CH_3 \cdot COOH$

Cleanliness the protection against harmful bacterial life. The necessity for a clean milk supply. — As the ripening of cheese depends on the bacterial and mold forms which change its structure and give it flavor, as the fresh sweetness of "June" butter comes from these organisms, so many acts essential to our comfort are carried on by these agents. The other side of the shield must not, however, be forgotton. Micro-organic life is after all destructive life. It is nature's way of breaking down and destroying the organic forms which have outlived their normal period of existence. Many of these micro-organisms are the scavengers of the world, bringing cleanliness, but bringing it through agencies of putrefaction and decay. Among the bacteria may be found some like the typhoid germ that causes typhoid fever, like the tubercle bacillus that is at the basis of consumption, and the pneumococcus, responsible for pneumonia. With these we are always at war. In consequence it is the housekeeper's duty to be always on the alert to keep her home clean. It is the citizen's duty to struggle for clean streets and clean meeting-places. By so doing we can guard ourselves from the attack of evil bacteria, and aid

THE MILK SUPPLY

Room in a modern dairy. Note the precautions taken to insure cleanliness.

and direct the action of the good. In no place is cleanliness more essential than in the milk supply. There, side by side with the cleanly, healthful lactic acid bacteria, are disease germs that kill thousands of children every month, every week, and every day. Many disease germs thrive and multiply in milk. There is one and only one safe way to keep them out — cleanliness. The farmer who keeps stable, cows, milk-room, bottles, and milk clean and sweet is taking another step towards saving the lives of the babies of the nation. A most necessary step this, for little children are dying daily in thousands from the dirty milk which is their only food.

CHAPTER XXVIII

FOODS

Food the source of man's energy. — The boundless wealth of energy locked up in the natural resources of this universe of ours, the stores of heat energy in fuel like wood and coal, of water power in lake or river, of electricity in magnetic revolving metal, are absolutely valueless in producing a state of civilization unless they are freed from their limitations by an energy of a higher quality than all the rest — the energy of man. Man's energy places the water wheel, constructs the steam-engine, digs from the bowels of the earth its hidden treasure, reaps from the soil the wealth of its production. To obtain his own energy man has but a single source of supply, his food.

Grain Elevators

Where the grain from the wide Western fields is stored.

No branch of the study of chemistry has more intimate connection with our daily life than have the problems which concern the three daily meals that provide the energy for every act of our daily life, the walk to the train or the game of tennis, the morning's lesson or the evening's pleasure.

The chemical composition of the foods. — Our foods are organic chemical compounds, substances which obey the same general laws that control the other compounds we have studied, though most of them are more complex than many of the inorganic bodies we have met. All of them, however, are made up from a very few of the common elements, constructed from those same primary units or bricks which make the matter of the world. In an article written on "The Food of the City Worker," the writer once discussed the relation which the elements found in food bear to the structure of the body, in a way which may bear repetition here:

"On the rear shelves of one of our old museums reposed for years a series of cubes and cylinders supposed to represent the proportions in which the various elements occur in the body of a human being of average size. I can feel to-day the strange and somewhat awful fascination by which those larger cubes marked nitrogen, oxygen, carbon, and hydrogen, those smaller cubes, marked phosphorus, sulphur, chlorine, iodine, sodium, potassium, calcium, magnesium, and iron, dragged one small boy's footsteps to their solemn shelf despite the attraction of full-rigged models of Chinese junks and armored Indian warriors. These cubes showed the essential parts of the body's frame, the lifeless elements, which, by the magic of living processes, build up the sentient organism, man. Since man is a compound of these elements, his food, from which alone can come the structure of the body, must be made up of these same substances. Since carbon, hydrogen, oxygen, and nitrogen chiefly constitute the body, our foods must be made chiefly from these elements. Since phosphorus and calcium in the form of calcium phosphate are largely found in the bones, our foods must contain both calcium and phosphorus to supply and repair bone growth and waste. Since life processes are continually breaking down the body structure, our diet must necessarily give the needed material for repair. We require something beside all this, however. We must have foods whose burning in the body will give heat, for heat gives two things needed for the proper running of the machine, power and a constant temperature well above that of the ordinary surrounding atmosphere.

The five classes of nutrients. — The meal which provided power for your morning labor or evening's pleasure, however diversified it may have seemed, was wholly composed of five simple classes of nutrients or substances which nourish the body. Though some students divide the foods into more divisions, the five classes of nutrients which we shall consider here are the proteids (sometimes called protein), the carbohydrates and fats, the inorganic salts and water. Each does its special part of the whole work, either by producing tissue and energy or by facilitating the processes through which other members of the nutrients are made available. Whether found in bread or cheese, meat or lettuce, each individual of these classes acts on the body in very much the same way, and the relation of food to the efficiency of the individual must be considered in terms of those basic groups rather than in terms of the foods made up by them.

Energy, heat, and repair in the human body. — In this quotation the emphasis was laid upon three factors, first, the energy which produces bodily power, second, the same energy turned to heat to keep up the body warmth, and third, material which shall provide tissue for the growth of the child, and for the replacing of those portions of man's structure which, as they are constantly broken down by the every-day acts of our life, need constant replenishment. Suppose we discuss these three necessities. We know that carbon in forming carbon dioxide yields vast store of heat, and that hydrogen oxidizing to hydrogen oxide does the same. When coal gas or water gas, alcohol or methane burn, these fuels, which are made up wholly of carbon and hydrogen, give off, as the products of their combustion, gaseous carbon dioxide gas and gaseous water. The foods which provide most of the power and heat furnished the body are the carbohydrates and fats, made up of the same elements as the fuels mentioned above, that is, of carbon, hydrogen, and oxygen. The burning of the first two of these elements in the oxygen of the air (supplied by our body bellows, the lungs) gives the energy which provides power for mental and muscular effort, and provides heat which keeps the body at that comparatively high temperature at which its processes can be most successfully carried on.

The carbohydrates. — Generally speaking, the carbohydrates make up the chief solids of the cereal foods, of wheat, corn, barley, rye, and oats. They are the chief nutrients of vegetables, like potatoes, beets, turnips, and squash, and of fruits, like oranges, bananas, and apples. The starch [$(C_6H_{10}O_5)_n$)] of the potato and the sugar of the sugar-cane ($C_{12}H_{22}O_{11}$) represent two of the most important classes of the carbohydrates. The carbon which they contain gives the chief fuel value of these foods, and in their case, as in the case of the other carbohydrates, the burning of the carbon and hydrogen provides most of the energy which the body requires. The "n"

SUGAR

A Cuban sugar factory, where some of the sugar for our meals receives its first preparation.

in the formula of the starch indicates a complex structure not wholly understood.

The fats. — The fats are the oily foods, like butter and olive oil. They are found in the fats of meats like bacon, and in the oily part of milk, which rises to the top of the pitcher in light cream. Like the carbohydrates these nutrients are composed of carbon, hydrogen, and oxygen, and the energy which they supply comes from the burning of the fuel ingredients which compose them. Containing a higher percentage of readily combustible matter than the other foods, they supply more heat and more energy in a given time. Thence comes the fat blubber diet of the Esquimaux, who needs quick heat-producing foods to enable him to survive the Arctic cold. A similar necessity prescribes cod liver oil and other fatty foods to the weak and old. The fats, like butter and olive oil, contain glycerine, $C_3H_8O_3$, and fatty acids like butyric, palmitic, and stearic. Their formulæ are commonly too complex for use here.

ORIENTAL HARVESTING

Functions of the carbohydrates and fats. — Much as these two great classes of food, which we have just discussed, can do, they cannot supply the main essential of body structure. They cannot build bones or muscle, brain cells or body cells. When fats and carbohydrates are taken in excess of the amount actually burned each day, the amount left over, so to speak, may be turned into fat and stored up in the body for future use. In time of limited diet or partial starvation these stores of fat can be drawn against for energy and heat to keep the body going.

The mineral salts. — Fat is but one of the needs of the body. Bone structure and cell structure are quite as important. The skeleton, that frame which holds up our moving, living parts, is largely made up of mineral salts, the third division of the foods, which, especially in their content of calcium and phosphorus, provide the bony solid structure which enables us to stand, sit, or move. Occurring in the smallest proportions of any of the five great divisions, these nutrients are as necessary as any. A deficiency in the amount supplied means defective bone structure and retarded organic processes. Salt, which assists in the carrying out of many body processes, is one of the mineral salts.

Occidental Harvesting

Water, and its place in man's diet. — Before passing on to the proteids, the last and in some respects the most important division of the foods, a word should be spoken of the importance of water in the daily diet. Man's body is composed largely of water, and the average man loses a large amount of water daily. That water must be replaced. It is necessary to the proper action of the blood, — that wonderful interior railway which bears raw materials and finished products of the body processes from point to point all over the human manufacturing plant. Many a life process in which individual organs are engaged requires water, as the individual manufacturing processes of a mill require the same liquid for washing, dissolving, or carrying substances in the natural course of manufacture. The necessity for the inclusion of water in the five great divisions of foods has been mentioned before in the chapter on water: "Plenty of water inside and out is one of the first rules of health."

The body cells. — The proteids, though partly used for the production of heat, are chiefly engaged in body building and repair. They supply cell structure, and the necessity for proteids will be recognized when we realize that every organ, every working part of the human body, is made up of cells. These are generally tiny complete parts of body structure, all composed of two and only two parts, a small central body, the nucleus, and a surrounding mass, the cytoplasm. Of the work of those cells Hough and Sedgwick have said in their "Human Mechanism," page 33: "The muscle and the gland (of man) consist of cells, just as the branches of the military service, the infantry, the cavalry, the artillery, the engineers, etc., consist of men. The cell is the anatomical or fundamental unit of these organs, as the soldier is the fundamental or anatomical unit of the army; in both cases the anatomical units, differing in equipment and training, perform different kinds of work, yet have the same essential structure, and the cells are combined into brigades, divisions, or corps, as tissues and organs; they make of the body an army organized to fight its way through the vicissitudes and against the obstacles of life." Since living organisms depend on these cells for life itself, their care and repair must obviously be one of the most important functions of the food.

The proteids. — Day in and day out, hour after hour, these living cells are constantly being destroyed, constantly being broken down by our varied exertions. It is necessary for the body's life that they should be replenished. The cells contain nitrogen and sometimes sulphur as essential parts of their structure. Neither carbohydrates nor fat, mineral salts nor water, contain these two elements. They must be furnished by the repair stuff of the body, the proteid division of the foods. The lean of meat and of fish, and the white of egg, are the forms in which we generally recognize proteid most readily, though a considerable percentage of this division of the foods is found in the breadstuffs, and quite an amount in milk. Proteid is the body-builder, the only nutrient which can provide the living structure of the human frame. Excess of this nutrient beyond the amount required for replacement may be burned to supply energy, power, and warmth, so that proteid can take care of some of the functions of the non-proteid part, of our food. The reverse action can never occur. The non-proteids do not possess the elements needed for cell building. They can never take the place of proteid. These foods are called the nitrogenous nutrients from the element which chiefly characterizes them, though they contain carbon, hydrogen, oxygen, and sometimes other elements as well.

The nutrients present in a few of our common foods. — Stating a few of the common foods in rough percentages of these nutrients we find that most of our foods contain a certain amount of every class we have mentioned. The percentage present in some of the common foods will be found below.

Bread has 37 per cent of water, 8 per cent proteid, 50 per cent carbohydrates, 1 per cent fats, and 2 per cent salts; milk has 87 per cent of water, 3.5 per cent of proteid, 4.75 per cent of carbohydrates, 4 per cent of fat, and .75 per cent of salts. Lean beef contains 72 per cent of water, 19 per cent of proteid, 3 per cent of fats, and 1 per cent salts, while peas contain 15 per cent water, 23 per cent proteid, 57 per cent carbohydrates, 2 per cent fats, and 2 per cent salts.

Soap

The making of soap. — The fatty compounds similar to those found in meat fats, in butter and olive oil, have still another use, the making of soap. The sulphuric acid produced in a country has been called a measure of its commercial advancement. The soap used by a nation has been said to be a measure of its civilization. For centuries soap has been made in the home by treating the fats saved from household processes ("the soap fat") with lye taken from wood ashes. The huge soap factories of to-day depend upon the same chemical action that went on in the kitchen of the home. The lye that was used in the big soap kettle on the fire was used for neutralization, and the basis of all soap-making in the manufactory or in the kitchen rests upon such action as takes place when fatty compounds (like the fat of meat or the oils pressed out of vegetables) are boiled with alkalies such as potash, or caustic soda (sodium hydroxide).

The chemical action which goes on in soap-making. — The fats which are used in soap-making are composed largely of two organic acids, palmitic and stearic, combined with an organic radical, glyceryl. The hydroxide or alcohol of this radical is glycerin. When you put fat in a kettle with an alkali and boil the mass, glycerin forms and separates out from the organic acids. That leaves these liquids free to act upon the caustic soda base which is present, and neutralization occurs. The palmitic acid forms sodium palmitate when neutralized with sodium hydroxide. The stearic acid forms sodium stearate with the same alkali. Soap made in this fashion is a mixture of sodium stearate and sodium palmitate. When potassium hydroxide is used instead of caustic soda, potassium stearate and potassium palmitate are in the soap. If other organic acids are present, such as oleic acid, some salt of those acids would be formed. In the case cited the salt would be an oleate. Soaps when made by this reaction are in solution. To bring them out of solution salt is added, whereupon the soaps present rise to the top of the liquid in a solid mass. The separation of soap as this insoluble substance is due to the fact that the sodium salts of these fatty acids are insoluble in brine. The reaction for the action between sodium hydroxide and the stearin of the soap may be written as follows: Stearin + sodium hydroxide = sodium stearate + glycerin.

$$C_3H_5(C_{17}H_{35}\cdot CO_2)_3 + 3NaOH = 3(C_{17}H_{35}\cdot CO_2)Na + C_3H_5(OH)_3.$$

This reaction becomes clearer when we realize that the radical $(C_{17}H_{35}\cdot CO_2)$, despite its many atoms, acts all together as a single atom would do. It simply exchanges a single trivalent radical (C_3H_5) for three atoms of monovalent sodium.

Various types of soap. — The type of soap which is produced depends on the fats and alkalies which are used in the making. When potassium hydroxide is the alkali a soft soap is formed. When castor or olive oil is treated with caustic soda a type of castile soap is manufactured.

The cleansing action of soap. — The cleansing action of soap depends on its chemical properties quite as much as does its preparation. When, after rubbing a lather of soap on our hands, we plunge them into water, a partial reverse action to the original combination takes place, and some of the caustic soda which was originally present forms again. This alkali cleanses off the oily matters of the skin, while the other portions of the soap, with the help of the water, remove dust and dirt mechanically.

Some of the street-car advertisements which state that a generous use of soap is the most important step on the road to self-respect are not far wrong. The lessons of the city baths and the tendency to civic cleanliness in our communities have proven their right to a place among the greatest factors in the uplifting of the people.

CHAPTER XXIX

THE CLOTHES THAT WE WEAR

The fibers of cloth. — Fibers are the beginning of our clothes. If you look through the microscope at a silk thread spun by a silkworm, at a fluff of wool sheared from the fleece of a sheep, at the white blossom of the cotton plant or at the fiber of the flax, you will find that each and all are filaments or cords, tiny tubes or cylinders of such a structure that they can be woven together. Indeed the threads which show along the ragged edge of any piece of cloth are made up from the winding together of small natural threads, the animal or vegetable fibers.

Vegetable and animal fibers. — Fibers vary considerably in their composition. Those which come from vegetable matter are plant cells of simple structure, generally some one part of the plant. Vegetable fibers are made up in the main of cellulose, a substance containing carbon, hydrogen, and oxygen, which is at the basis of all woody fiber and which is the principal constituent of wood. The fibrous properties of cellulose have already been mentioned in the chapter which told of the making of paper. Animal fibers, on the other hand, are bodies which generally contain nitrogen and often sulphur. In composition they are very much like the proteid matter which forms so great a part of our food. Under the microscope animal fibers are seen to be either complex cell structures (in this case bundles of cells tied up as it were in a single envelope) or else cylindrical filaments. A typical animal fiber, silk, is produced by the silkworm in much the same way as the spider spins its thread, or the caterpillar its cocoon, that is, from a secreted liquid.

Cottonseed oil and linseed oil. — Both cotton and flax are grown for the oil compressed from their seed as well as for their fiber. Cottonseed oil is used in making vegetable substitutes for lard, such as cottolene, in the preparation of soap, and in the making of cheap grades of salad oil. Linseed oil, the oil extracted from the seed of the flax plant, is far more important than cottonseed oil as it is the most widely used of the so-called "drying" oils, a group of great value to painters. These oils have the property of absorbing oxygen upon exposure to the air (as when spread with paint upon the side of a building) and changing from a liquid oil into a dry, hard varnish. This is a different action from the evaporation of alcohol in varnishes. The change is accelerated by boiling the raw oil with certain chemicals called "driers." The press cake left behind after the oil has been pressed out from the flaxseed or cottonseed is a valuable cattle food. So much for the uses of the seeds. The fibers of the plants are the parts which interest us most.

COTTON
The cotton-press yard at New Orleans.

Cotton and flax fibers. — Vegetable fibers are divided into seed hairs (single cells like the fiber of cotton) and bast fibers (bundles of cells joined together, we may almost say glued together, like the fiber of flax which forms linen). In the growth of the cotton plant the fibers at first cover the seed with a hairy envelope, then, as the plant develops, a mass of lint is formed which breaks the cotton-pod or boll. After the boll bursts, the fiber, acted upon by sun and air, twists into a spiral. The fibers of flax show strongly marked characteristics under the magnifying glass. Each individual cell is a long cylinder pointed at each end, thick-walled, but with a hollow center like a gas or water-pipe. These cells are from two to four centimeters long and are marked with crossing bands. The individual fibers are glued together in bundles which look somewhat like the bound rods of the Roman fasces, those symbols of the republic's authority.

The preparation of flax fibers. — Temperate zones produce the best flax for textile use, and the large amount of the plant which is grown in tropical or sub-tropical regions is cultivated chiefly for its seed. When its fiber is desired the plant is pulled up by the roots before the seeds ripen. It is then "rippled," that is, the seeds are detached, and afterwards "retted," soaked in water or some other liquid to break the fiber from surrounding woody tissue. The preparation of the fibers of hemp and jute resemble in a general way the preparation of flax fibers.

The silkworm and silk thread. — The silkworm produces silk merely as one part of its life process. The worms are now largely raised in incubators where they are kept at a temperature of from 18 degrees to 25 degrees C. during their growth. The caterpillar of the silkworm has one of the largest appetites on record. It feeds voraciously on mulberry leaves for some thirty days, growing rapidly and finally reaching a length of about three inches. Once it has attained full growth its appetite ceases, the caterpillar stops eating and crawls upon twigs where it begins to spin. For three days it spins, forming a cocoon quite similar to the cocoons you may find in fall lying on the veranda or attached to dead leaves. During the spinning the caterpillar unceasingly produces two liquid threads which solidify and join together as they reach the air. The cocoon when finished is made up of one continuous fiber from one thousand to four thousand feet long. This continuous fiber is taken off on reels directly from the cocoon, and is subjected to a series of processes intended to get rid of its harsh feeling, and its stiff, coarse appearance. By various boilings and soakings the thread is brought to a soft, glossy condition.

SILK
Dyeing and winding silk in China.

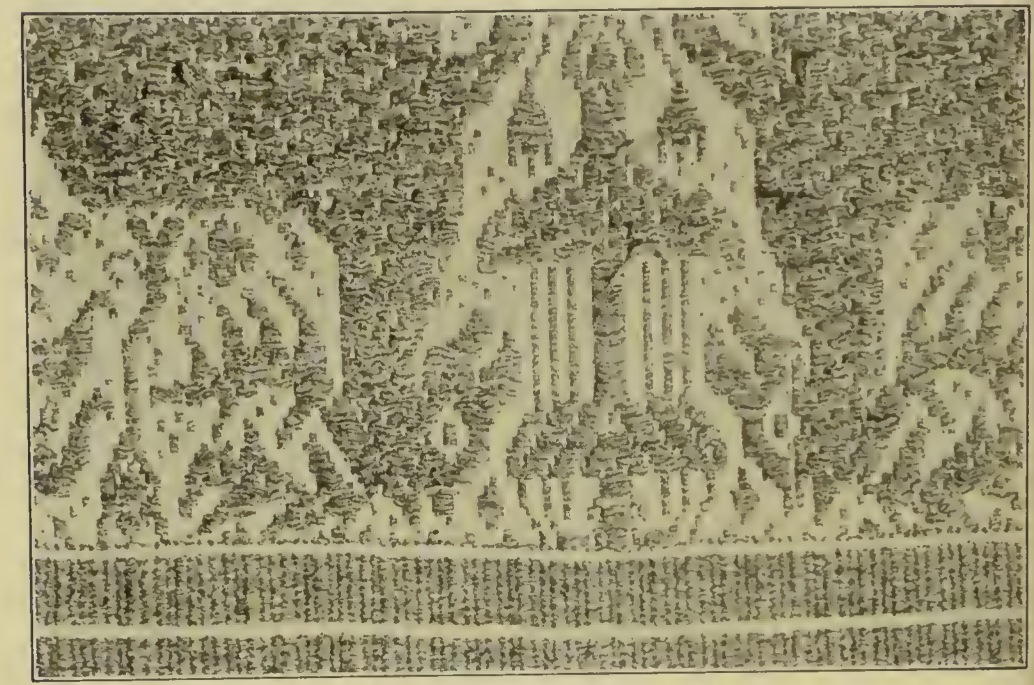
Wool. — The manufacture of silk has always been a process a little apart from our common knowledge, but the production of woolens, from the first step of sheep-shearing through the spinning of woolen thread and the final manufacture of home-spun cloth, was the every-day affair of our ancestors a generation or so ago. Wool, it is true, is a general name which beside the fleece of sheep includes mohair, cashmere, vicuna, and alpaca threads which come from the hairs of different species of goats. All the other sources of animal fibers put together, however, give scarcely a fraction of the textiles that come from the back of the sheep. Wool is very much like ordinary hair in most of its characteristics and differs mainly in its physical structure. Each fiber of wool resembles one of those queer-shingled Japanese towers with overhanging eaves, for every fiber is covered with a layer of shingle-like scales, all projecting in the same direction, and all overlapping like shingles on a roof. The length of the fiber may range from one to ten inches or even more. The longer, coarser wool is generally woven into worsted goods. The shorter, finer wool makes woolen goods. Wool, like silk, follows the habit, of many proteid substances and contains nitrogen and sulphur in addition to the three elements which range through the whole of organic chemistry, carbon, hydrogen, and oxygen.

WOOL
Sheep ready for shearing.

The preparation of threads. — After the wool is sheared from the sheep it is freed from dirt, grease, and impurities by washing and scouring, but the threads, be they cotton or flax, wool or silk, are not ready for use even when they have been quite thoroughly cleaned. No natural fiber is free from coloring matters, impurities, or other foreign bodies, which would either injure the cloth directly or keep dyestuffs from exerting their proper effects. For this reason, beside the washing and scouring, all natural fibers must undergo a bleaching or chemical cleansing.

The bleaching of cotton cloth. — Cotton yarn, after being washed with boiling soda ash, is bleached with chlorine generated from bleaching powder. (For bleaching with chlorine see chapter on chlorine.) Cotton cloth is boiled with various solutions containing caustic soda, soda ash, hydrochloric or sulphuric acid, and bleaching powder, according as the goods is to be bleached for use as white muslin, for use with the red dye formed by alizarin (the old Turkey red bleach) or with the so-called madder bleach which is used for those cottons which are to be dyed with various colors.

Linen bleaching. — Linen bleaching resembles cotton bleaching, in its essential methods, but adds another well-known household process. Unbleached linen is commonly whitened by being spread out on the grass in the sunlight, where it will be wet with dew or rain, and exposed to the action of the light and air. The ozone in the air is believed to be the chemical that does the bleaching. Here as in most bleaching processes the action does not work well when the material is dry. It works better when the textile is regularly dampened.



LINEN
Old Italian towelling.

Wool and silk bleaching. — Animal fibers cannot be bleached with chlorine or the other agents employed in bleaching vegetable fibers, as they would be injured by such comparatively harsh treatment. In consequence fibers of this kind are very thoroughly scoured and afterwards bleached, either by exposing the yarns on wooden rods to the action of hot sulphur dioxide gas (the mild bleacher) (for bleaching with sulphur dioxide see chapter on sulphur dioxide) or to the action of sulphurous acid (the solution of sulphur dioxide in water). Hydrogen dioxide and sodium bisulphite are both used in bleaching wool, while in silk, as it is commonly a more costly product, the manufacturers use even more of the expensive bleaching agents. Hydrogen dioxide, potassium permanganate, sulphurous acid and sulphur dioxide gas, all these are employed for silk bleaching.

Coal-tar dyes. — Modern dyeing is a very recent art whose history goes back only a trace over fifty years. When you read stories laid in the times of the crusades or in the days of Rome which tell of pageants filled with color, or of great tournaments where the lists shone with a thousand tints, you can know that the author of such tales knows nothing about the history of dyeing. Up to 1856, when Sir William Henry Perkin discovered the first aniline dye, the first dye to be produced from coal tar, there was but the slightest range of choice, and the selection of a dress or a cravat was confined to blacks and whites, grays and indigo blues, turkey reds, one or two purples, one or two greens, and just a few combinations produced with these. Then came the beginning of the revolution that has changed the world of fabrics from a dull-hued, somber desert to a brilliant, tropical garden. Perkin, a boy of eighteen, seeking to make quinine, separated out a colored substance from coal tar, which was the first coal-tar dye, mauve. After overcoming tremendous obstacles he established the possibility of dyeing with similar dyes. From Perkin's production of the dye mauve came the thousands on thousands of shades in use to-day. Rainbow colors in the greatest profusion to beautify cloth and ribbon are drawn from the thick black liquid coal tar left from coal-gas making. On Perkin's discovery rests the dyeing of textiles the world over.

Dyeing. — We have already seen that the preparation of substances made into textiles depends on their fibrous nature. Their dyeing depends on fibers quite as much. Dyeing in general requires the fixation of some colored substances inside the fiber. This substance must be put there in such a way that it will not wash out or, to state it otherwise, a fabric is dyed when some insoluble colored compound has been fixed within the fiber. The method of depositing that insoluble substance inside the fiber was long kept a mystery of mysteries by the dyer. It is an open secret to-day.

Direct dyeing of animal fibers. — Animal fibers like silk and wool are constituted quite differently from vegetable fibers like cotton and linen. Silk fibers are solid. Both silk and wool contain, naturally, certain chemicals which make it possible for them to be dyed directly. If we dip white silk into picric acid, one of the common organic dyes, we take the fabric out dyed a soft yellow. This yellow is permanent. It does not wash out. The picric acid has united with a natural chemical which permeates every fiber of the silk and has formed an insoluble compound in combination with it. If we dip wool (the other fabric woven from animal fibers) into picric acid we find that this also can be dyed directly. As in the case of silk some natural compound exists within the fiber, which precipitates the added dye, forming a permanent insoluble colored compound which colors the whole fabric.

Why vegetable fibers cannot be directly dyed. — When we dip cotton or linen (vegetable fibers) into picric acid we obtain no such result. A temporary color is shown on the outside of the goods, but this rapidly washes off leaving the fabric as before. This is true in general of all vegetable fibers. They cannot be dyed directly because they possess no substance within their fibers which will combine with any single dye. Two things, the fact just stated, and the second fact, that those dyes which will dye silk and wool directly are few in number, bring about the necessity for a second more general method of dyeing.

Indirect dyeing. Mordant, dye, and lake. — In this second method a dye which would of itself wash out of the fibers is converted into an insoluble compound by the addition of another substance. We could illustrate this action by taking a glass tube to represent the tubular fiber. Now if we fill this half full of a clear solution of silver nitrate we should have a substance which, though not a dye, would represent one for our purposes. If we then fill the tube with hydrochloric acid we should have an example of a precipitating solution. Insoluble, white, silver chloride would be precipitated when the two meet inside the tube, and the solid white particles formed would reflect back white light. If we could see what goes on within the fiber we should find that a not dissimilar action occurs in dyeing. A colored or colorless dye in solution within a tubular fiber is precipitated by another solution. A fixed insoluble substance which cannot escape is formed by the action. A substance which, added to a dye, fixes a colored insoluble compound in the material is called a mordant. A colored insoluble compound formed by the combination of a dye and a mordant is called a lake.

The range of indirect dyeing. — Since the fixing of a dyestuff within a fiber by mordanting is quite independent of the nature of the fiber, it is evident that here is a process which applies to the vegetable fibers, such as cotton and linen, which resist direct dyeing. The process applies also to animal fibers, such as silk and wool, which can be dyed directly, and the number of coal-tar colors which can be turned into soluble compounds inside of fibers ranges through almost every shade of every color. Since modern fashion demands the widest possible field of color choice, it is evident that dyeing must depend chiefly on this method.

The use of alizarin. Mordanting in general. — The use of the dye, alizarin, illustrates the action of dyes and mordants particularly well. If we drop an unmordanted piece of cloth into a water solution of alizarin, the cloth is colored a dull temporary yellow. This yellow disappears on washing. No dye has been fixed in the fiber. If, on the other hand, we drop into a bath of alizarin a cloth which has been previously mordanted with an aluminum compound we obtain a permanent red dye. If we mordant the cloth with a solution of iron salt we get a purple dye. If we use a tin salt we find that a pink dye is fixed within the fibers. The final dyeing depends on the interaction of the dye and mordant to form an insoluble lake within the fiber. Two actions, however, commonly take place in preparing the mordanted cloth. In the process of mordanting the first necessity is to expose all fibers of the fabric to the action of the mordant. In order to do this the cloth is generally soaked in the dissolved mordant until the fibers of the fabric are impregnated (the mordant is commonly a solution of a metallic salt). In the second place the mordant must be so secured in the fiber that it will not be washed out when the cloth is run into the dye-bath. That is, the mordant must be fixed in the fiber in such a fashion as to make sure it meets the dye. This can be accomplished by various methods. In one process a weak alkali, ammonia for example, is used, which forms a metallic hydroxide with the mordant. The mordant under these circumstances enters the bath an alkali, and reacting with the acid dye precipitates an insoluble salt. Occasionally these conditions are reversed and a basic dye reacts with an acid mordant. Cases of this sort are comparatively few. Some printed goods like calicoes are colored on one side only. Under such circumstances a solution is saturated with the mordant mixed with the dye and thickened with some foundation such as starch or gum. The pattern desired is printed on the cloth much as type is printed on paper. After the printing the cloth is put through a steaming process which leaves the dye fixed.

The names of a few dyes. — To give even a brief review of the many dyes which chemists have formed from coal tar would take many pages; only a few can be mentioned here:

malachite green, which gives a dark bluish green; rosaniline or fuchsine, which dyes all textiles a brilliant magenta; aniline blue, which forms a fine clear blue; bismarck brown, which imparts a dark, rich brown; indigo, which fixes the well-known indigo blue.

Two examples of the economic effect of chemical research. Alizarin. — Two of the great dyestuffs, alizarin and indigo, illustrate by their history the wide-spreading economic effects of chemical research. Alizarin is found in the madder plant, the "Rubia tinctoria" of the botanists, a tree whence came the naming of the wide-spreading order of plants, the Rubiaceæ, an order which includes the copper tree, cinchona, the tree from which comes quinine, the Sierra Leone peach and many other plants noted for beauty and fragrant blossoms. From the earliest times the extract of the madder has been valued for its dyeing powers. It was so widely used in the middle of the last century that whole provinces of continental Europe were given up to its cultivation, while the production and extraction of the dye from the plant was a business of very great importance. Sir William Henry Perkin's discovery of mauve came like a thunderbolt from a clear sky, and the world awakened to the possibility of swiftly extracting dyestuffs from coal tar instead of waiting for the slow processes of the soil. That discovery sounded the knell of the madder industry. Research followed research, and in 1868 Graebe and Liebermann made alizarin in the laboratory. Their process was not at first a commercial success, but other methods developed from that opening. To-day the whole of the alizarin used in commercial work is made in the laboratory. The thousands of acres formerly used for the cultivation of the madder have been turned to other uses, set free for beet-sugar raising, and for varied crops. The many men employed, the great wealth involved in the raising of the madder plant, and the extraction of its dyeing juice, have turned to other channels.

Indigo. — The story of indigo is similar. Known as a dye since the Picts stained their bodies with the blue coloring of the woad, this substance found in the indigo plant of the tropics and the woad of the north was produced for years on huge indigo plantations. Even after many other dyes had been produced in the laboratory, artificial indigo remained undiscovered and the indigo planters went on reaping large profits. In 1880 the end of natural indigo began. In that year Baeyer prepared indigo blue synthetically, building it up artificially in the laboratory. As in the case of alizarin the manufacture of indigo was not at first a commercial success, but not many years passed before methods were discovered by which the dye could be made at a price much less than it could be manufactured from the plant. Those discoveries brought about the doom of the old indigo trade, and synthetic indigo took the place of the dye made in nature's laboratories. Once more science had improved on the processes of nature.

CHAPTER XXX

HEAT AND LIGHT IN CHEMISTRY

Endothermic and exothermic reactions. — Two experiments which illustrate the relation of heat to every-day experimentation are carried on in thousands of kitchens throughout the land. Whenever coal burns we have oxidation accompanied by the evolution of heat. Whenever ice and salt are placed together in an ice-cream freezer, we have solution accompanied by an absorption of heat from the can of the freezer and from the atmosphere around. Among the general characteristics which persist through the whole body of collected chemical knowledge, few are more noticeable than the heat phenomena which accompany chemical action. Such phenomena may be divided into two classes. One in which heat is absorbed, a second in which heat is given off. Most of the reactions we have considered may be placed in one of these two groups, the first group called endothermic or heat-absorbing; the second group called exothermic or heat-developing.

The calorie. — All types of chemical reactions which involve heat are measured by the heat given off or absorbed. A unit of heat known as the calorie is used for such measurement. The calorie is defined as the amount of heat required to raise one gram of water one degree Centigrade. Suppose you had a dish containing one gram of water placed over a gas-burner. If we light the burner and (disregarding the amount of heat which goes into the dish) heat the water just enough to bring its temperature up one degree so that its temperature rises for instance from 4 degrees to 5 degrees C., one calorie will have been used in heating one gram of water one degree. If we heated one hundred grams of water till the temperature rose one degree, we should use 100 calories. If we heated one gram of water through one hundred degrees we should also use 100 calories. If we heated twenty grams of water through forty degrees we should use 800 calories.

Specific heat. — A gram of water requires more heat to raise it through one degree than any other common body. If we place ten grams of iron and ten grams of water on the same stove at the same time the iron will reach 100 degrees C. long before the water reaches the same point. Every body requires a definite fraction of a calorie to raise it one degree, and the number which represents this fraction is called the specific heat of the body. To state it more exactly, specific heat is the amount of heat required to raise one gram of any substance one degree Centigrade, compared with the amount of heat required to raise one gram of water one degree Centigrade. A substance which has a specific heat of .2, therefore, requires only one fifth the heat that water requires to produce the same effect. One fifth of a calorie would do as much for a gram of this body as a whole calorie does for water. A whole calorie would raise a gram of such a substance five degrees.

The synthesis of water as an exothermic reaction. — When hydrogen and oxygen unite to form water a very large amount of energy is liberated. Expressed in calories the amount of heat evolved in forming eighteen grams of water is 68,400. calories. In this case sufficient energy is given off to raise one hundred grams of water to 684 degrees C. This is an exothermic reaction. Heat is given off.

The analysis of water as an endothermic reaction. — When we decompose water into its elements we find that a precisely similar amount of energy is required to decompose the compound. 68.400 calories decompose eighteen grams of water into the elements which compose it. Such an action is endothermic. Heat is absorbed. We could express this in reaction form as follows:

$$2 \text{ grams H} + 16 \text{ grams O} = 18 \text{ grams } H_2O + 68{,}400. \text{ calories.}$$
$$18 \text{ grams } H_2O = 2 \text{ grams H} + 16 \text{ grams O} - 68{,}400. \text{ calories.}$$

The addition sign (+) of 68.400 calories in the first case indicates an exothermic reaction. The subtraction sign (−) of the same amount in the second case indicates an endothermic reaction.

The law of the conservation of energy as applied to heat reactions. — The statement that the amount of energy given off in the combination of any substance always equals the amount of energy required for their decomposition is a corollary of the law of the conservation of energy. "Just so much energy as you put in, just so much energy you take out." In this special case a law has been formulated as follows:

"The amount of energy manifested in a reaction is equal to that manifested in the reverse reaction."

The calorific power of burning carbon. — Of all exothermic reactions the formation of carbon dioxide which releases 97.000 calories when twelve grams of carbon unite with thirty-two of oxygen is probably the most widely used and the most practically effective. On this reaction largely depends every type of ordinary household burning. Next to this in importance would probably come the oxidation of hydrogen to water, since so many of the fuels contain hydrogen which can be oxidized, as well as carbon.

Coal gas. — The composition of various fuels has already been discussed in the preceding pages. The production of one of the common laboratory fuels, coal gas, has so intimate a connection with the constitution of flame, that the account of its properties and manufacture has been left to this place.

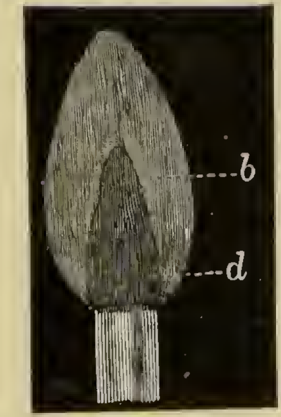
Coal tar, ammoniacal liquor, and coal gas. — Three substances distill over when coal is subjected to a destructive distillation, coal tar, a liquid known as ammoniacal liquor, and coal gas. The coal tar serves as a basis for dyes, medicines, and other synthetic laboratory products. The ammoniacal liquor furnishes household ammonia. The coal gas gives light and heat. More than ninety-five per cent of ordinary coal gas is composed of carbon, hydrogen, and oxygen. It commonly contains some nitrogen and some sulphur, as impurities. The greater part of the three elements first named is in the form of methane, hydrogen, and carbon monoxide. When they burn these gases furnish heat but little light. Other organic gases, like acetylene, chiefly furnish the illuminating power.

The manufacture of coal gas. — The part of a gas-plant which we usually recognize is the big round gas-holder that stands beside so many gas-houses. That gas-holder really belongs at the end of the process of manufacture, for it is from that point that gas is delivered to the pipes which run out through the city. The gas has been through a whole series of processes before it passes into the gas-holder. The first step in gas-making is the heating of the coal in retorts. From these retorts come forth the volatile products mentioned in the preceding paragraphs, and the various pieces of apparatus which intervene between the retorts and the gas-holder are all mainly intended for one purpose, the cleansing of the gas from impurities. After leaving the retorts the gas, mixed with ammoniacal liquor and coal tar, passes first through a hydraulic main, as it is called. There the gas bubbles through the water and a part of the impurities are left behind. This hydraulic main besides cleansing acts as a seal to prevent the return of gases to the retorts. The gas passes from this point into a series of long tubes which compose what is known as the condenser, where the gas is slowly cooled, and the liquor and tar which has passed by the hydraulic main is still further condensed. From the condenser the gas passes into the scrubber. This is filled with loose coke or stone on which a liquid trickling down meets the gas coming up, and dissolves, or holds in suspension, still more of the impurities. The purifiers come next. These are trays covered with lime which intercept carbon dioxide and sulphur compounds that may be present. On leaving the purifiers, the gas passes into the gas-tanks through a meter which records the volume of the passing gas. The pressure necessary to keep the gas in movement and to balance the tank is largely furnished by an engine called an exhauster. The gas-tank itself is made much like an ordinary round butter-box turned upside down. The upper part floats in the lower part, which is partly full of water. This top rises as gas enters, falls when gas passes out, and keeps an even pressure on the gas in the city mains aided by weights and pressure from the exhauster. The gas is kept from escaping from the gas-tank by means of the water seal in which the upper part of the tank rests.

Flame

The ignition-point. — The construction of a flame which rises from a lighted gas-burner, like that which springs up from flaming wood or coal, is of direct importance to the welfare of man. Every time a Bunsen burner is lighted, every time a meal is cooked, the construction of the flame and the way in which it is applied determine to a marked degree the effectiveness of the process. Few signs indicate the practical interest taken in fuels to-day more clearly than the fact that nearly all great corporations regularly test all fuel which they receive in order to make sure that they are getting the greatest possible result for the money invested. Substances which can be used as fuels burn when their temperature is raised above the ignition or kindling point. This temperature differs with each individual substance. With some bodies, like phosphorus, it is very low, with some, like the metals, it is very high. In ordinary exothermic combustions the ignition-point of a substance is lower than the temperature produced by its burning.

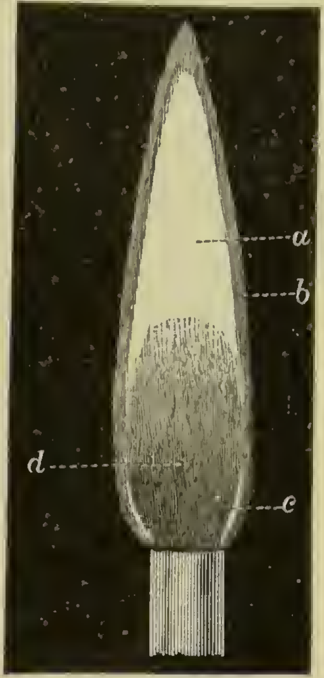
The structure of a flame. — Flame occurs in combustion only when both the substances combining are in the form of gases or vapors. The wax in a wax candle, the sulphur of the sulphur match, the coal gas in the burner, all are either gases at the beginning of the combination, or may be vaporized when heated to the ignition-point. When hydrogen burns, forming hydrogen oxide (water), we have the simplest possible case of two gases uniting with the formation of flame. When such a flame as this appears at the end of a tube like that of the ordinary gas-pipe or Bunsen burner, its interior portion is a hollow cone which consists of unburned gas. The flaming portion is a cone which surrounds the hollow part. No combustion takes place in this hollow interior cone, a fact which may be easily observed by thrusting an unburned match into the unburned gas and observing that the phosphorus does not burst into flame. The gas must meet oxygen before it can combine. Since it cannot obtain this oxygen in a flat layer in the Bunsen burner it springs out into a cone-shape. In an ordinary coal gas flame or in a candle flame we find a burning similar to the oxidation of methane and ethane. In methane and ethane the heat of the flame is due to the combining of their carbon and hydrogen with the oxygen of the air. With complete combustion all the carbon of these substances finally oxidizes to carbon dioxide, all their hydrogen forms hydrogen oxide (water).



FLAME. I

Construction of a hydrogen flame. *d* is the hollow, inner space of unburned hydrogen. *b* is the actual burning portion.

What happens when air is admitted at the base of a flame. — In these flames as in hydrogen the hollow space which consists of heated unburned gas appears. When the luminous flame is used the flame shows three parts above this space: a slightly luminous outside portion, a brightly luminous middle mantle, and a blue non-luminous region towards the base of the flame. When we admit air to the base of the flame, as in a Bunsen burner, where we admit air through a hole at the base of the tube, a new condition appears. A slight amount of air reduces the size of the luminous portion. Add still more air and we reduce it still further, while if we admit a sufficient amount of air the luminosity disappears altogether, leaving two parts to the flame instead of three, the blue region, which is now an inside cone, and an outer cover of flame. The middle portion of the three parts under these conditions disappears. This disappearance is due to the change in the chemical reactions which go on. When the air is admitted to the flame, the cooling effect of this air and the oxygen passing through so affect the flame that its luminosity ceases and the luminous gases which give light are no longer formed. In consequence, with a sufficient addition of air the constitution of the flame is limited to those gases which burn without evolution of light. For many years it was believed that the light properties of the flame depended wholly on the presence of slight particles of solid matter in the flame, or, to speak specifically, that the flame of the gas-burner resulted from the unburned carbon present, which, heated to incandescence, gave light. To-day other theories have been brought forward. According to some recent authorities the reason why the flame gives light may be due either to the density of the flame gases, or to the introduction into the flame of solid matter. According to yet other authorities, the Bunsen burner flame owes its non-luminous powers to oxidation, dilution, and cooling. The striking back of a burner flame (the common laboratory occurrence where the flame lights at the base of the burner instead of at the top) is due to air, so altering the usual amounts of gas and air which ascend the tube as to form an explosive mixture. When this explosive mixture is formed the flame at the top sets fire to the inflammable mixture within. The flame of this mixture running down the tube lights the gas as it emerges from the opening in the base.



FLAME. II

The three parts of a flame. *a*, the bright, yellow luminous region. *b*, the faintly luminous mantle. *c*, the blue region at the base. All the foregoing are a part of the actual flame ; all are burning. *d*, is the dark region which consists of unburnt gases.

The oxidizing and reducing flames. — The two flames of the Bunsen burner are sometimes spoken of as the oxidizing and reducing flames. The outer area contains oxygen in excess, the inner area contains heated and unburned hydrogen or hydrocarbon. The action of these two parts of the flame may be clearly illustrated in the following fashion: if we let a piece of bright copper wire cross a Bunsen flame the heated outer edges will turn to copper oxide. The portion which rests on the inner cone will remain bright. If we now move the oxidized portion into the inner flame, the copper oxide will be reduced to the metallic form and the bright metal in it so disappears. The hydrogen or the hydrocarbon within, seeking for oxygen, has taken it from the oxide of the copper.

Flame. III

A luminous mantle, similar to the Welsbach burner, which gives a bright white light when heated to the point of incandescence.

The electric furnace. — To turn from ordinary ways of heating to extraordinary methods, a word should be said here of the electric furnace, which will give the highest heat of any heating apparatus yet constructed by man.

The construction of Moissan's furnace. — This furnace in its essential is a very simple piece of apparatus. As used by Moissan, its inventor, it was nothing more or less than a box made of two blocks of quicklime placed one on top of the other and hollowed out inside. Quicklime was used because it is infusible at the highest temperature obtainable, and is, besides, a non-conductor. The hollow within the box contains the crucible in which rests the substance to be fused. An opening is cut through the blocks and two rods of carbon are thrust through, which meet just above the hollow in which the crucible rests. These rods do not touch each other, but allow room for an arc similar to the arc of the arc light. The space between the two terminals allows the electric current to spring from the positive pole to the negative pole. After the furnace has been set up in this fashion a current of enormous power is sent through the carbon electrodes. As the current jumps from the positive to the negative pole, the tremendous temperature of 3500 degrees C. and over is evolved. This great heat acts directly upon the crucible below. The furnace used by Moissan was about 1 foot by 8 inches by 1 foot in dimension.

Heat equations in daily life. — The great corporation measures the energy of its fuel in calories. The scientists, who recognize that the energy of the man whom the corporation employs comes from his food, figure food energy in calories, and modern students of dietary conditions largely base their beliefs in the comparative value of foods upon the fuel value which those foods render. In all our discussion of the theoretical side of the chemical changes which involve heat, the thermal changes which are nearest at hand should not be forgotten. Every act of our existence is dependent on the food which furnishes us with the energy of heat. That heat can be measured by calories.

Photography

Light reactions in chemistry. — The effect of sunlight in changing and fading the color of dress or ribbon, the combination of hydrogen and chlorine to form hydrochloric acid gas under the influence of light, the breaking down of the silver compounds on a photographic plate when exposed to a like influence, each and all illustrate the effect of light on chemical compounds. Why sunlight fades one color and not another we cannot tell, nor do we fairly understand such combinations under the influence of light as are shown in the example just cited of the formation of hydrochloric acid gas. Of the third chemical change effected by light, although we know little enough, at least we can give approximate reasons for the magic picture which the sun throws upon a photographic plate. Yet though we have some knowledge of the reasons for positives and negatives, even here we cannot yet tell the exact chemical processes which go on in a photographic compound which is attacked by light.

The effect of light on silver salts. — As a rule, light acts more powerfully upon the halogen salts of silver, that is, upon the bromide, chloride, or iodide of this precious metal. If you precipitate white silver chloride from a solution it gradually darkens to violet in the sun. A not dissimilar change takes place with other salts affected chemically by light, and colored compounds are formed whose composition has not been definitely determined. The change is, however, probably due in part to such a reduction of compounds as to make it possible by developing the plate or film to obtain silver in the metallic state. When light strikes silver chloride, for example, a part of the compound breaks down into its component parts. Where more light falls, more of the salt would decompose. If you were taking a photograph of a man who wore a dark coat, a white collar, and a light silk tie, the silver salt on the photographic plate would break down most completely where the light reflected from the white collar. It would be broken down a little less by the light from the light silk tie. It would break down less completely where the light from the man's face attacked it. It would break down least of all where it was affected by the light from the dark coat. The light has affected the plate so that developing produces a black spot of metallic silver at the collar, a lighter shade on the face, and the lightest stain of all on the coat. A picture thus taken is, therefore, a negative. The lighter the object, the more of the silver salt broken down, and as a result the darker the stain on the plate when the developer gets to work.

The making of a positive from a negative. — Every substance which is dark breaks down less of the silver salt, and therefore, on the negative, the dark trees are the lightest. To get the true shadings it is necessary to get a positive from the negative. This may be done by placing the negative on photographic paper and exposing the combination to the light. The dark spots of the negative, where dark metallic silver lies, shield the sensitive paper. The lighter spaces, which were but slightly affected by the original light, now allow the rays to pass easily. In consequence the positive is thin where the negative is thick, thick where the negative is thin. The original tone values have been restored and the picture shows shadings as they are.

The sub-salts of silver formed by the action of light. — The action of light on silver chloride or other photographic compounds is by no means confined to the production of metallic silver alone. Various compounds known as sub-salts, oxychlorides, and double compounds of silver are formed, all of which probably aid in one way or another in bringing about the effects produced.

Developing a photograph. — In the ordinary practice of photography, the photographic salt, whatever it may be, is held in a gelatine layer on a glass plate or a film. When the plate is exposed a negative forms in the fashion already explained, but before the negative can be seen it must be developed. This is done by the use of some reducing agent such as the old standard pyrogallic acid or ferrous sulphate. When the developer is poured over the plate the whitish film gradually darkens and little by little the image printed by the sun's rays appears. This image is largely a deposit of finely divided silver. Some of the silver salts which are present beside the silver and which have not been changed by the light would be clouded if the plate were taken from a dark room without further treatment than that given it by the developer. In consequence these silver salts must be removed. This is done by fixing, as it is called, that is, washing off the silver salts with sodium thiosulphite solution commonly known as "hyposulphite." After the final cleansing by water and drying, the negative is then ready for use. Most prints must be fixed like plates. Toning by use of the solution of gold or platinum salts is commonly employed to give a finishing touch.

CHAPTER XXXI

IRON AND COPPER

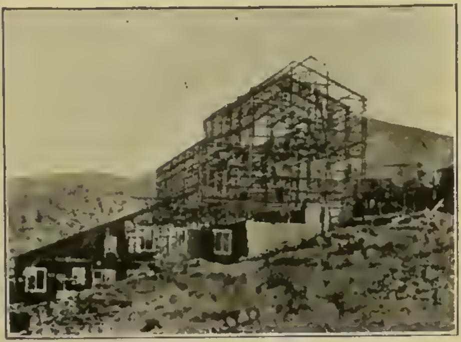
Symbol for iron = Fe. Symbol for copper = Cu

The reduction of iron ore to iron. — When some primeval smith first heated red, soft, iron-bearing earth with glowing carbon, and, to his amaze, wrought out an iron axe, he blazed a first great step upon the road that leads to our to-day. The present world is spider-webbed with iron. Iron railway engines drag iron cars on iron rails through cities filled with iron machinery and built by iron tools. The seas are crossed with

THE FORGE

iron liners driven by iron engines. Even the air is now traversed with ships whose motor power rests firm on filament-like skeletons of steel. All of these things have come because some prehistoric smith found by chance or study that if he heated iron ore (iron oxide) with carbon he could obtain metallic iron.

The strength of iron. — No other metal offers service to the world like this. Gold and silver, lead and mercury, could wholly disappear and not leave a fraction of the gap that iron's loss would open. Its dominance is inevitable, resulting from the fact that beyond all other metals it possesses properties useful to man. Suppose we list a few of them here.



Iron the Builder. I
The steel frame of a modern building.

Iron the Builder. II
Same building completed.

First: Taking it foot for foot, a metal beam in a floor or a metallic piston in the cylinder of a steam-engine will bear approximately twice the weight when made of iron that it will bear when formed from any other metal. Iron has double the strength or tenacity possessed by any other common metal.

The lightness of iron. — Second: The lightness of iron makes it available for ships, towers, high buildings, and many other structures where light weight and strength need to be combined. Iron is one of the lightest common metals; it is about a third the weight of platinum, a little less than a third the weight of gold, about half the weight of mercury, and four fifths the weight of copper.

The rigidity of iron. — Third: The rigid support which keeps the propeller blades of a steamer unshaken by a transatlantic gale, or enables huge pillars to bear the weight of a twenty-story building, is given best by iron. Iron is one of the least yielding or malleable of the metals.

The infusibility of iron. — Fourth: Because of the infusibility of the metal (its capacity of holding fire without melting) all our houses are warmed, our dinners cooked, and our water boiled in iron furnaces, stoves, or boilers. Iron is the least fusible of ordinary metals. That makes it excellent for fireproof building material.

Iron and carbon. — The percentage of carbon in iron determines a division of the product into three great groups, each of which is particularly available for some special task. These three divisions are wrought iron, steel, and cast iron.

Wrought iron. — The glowing horseshoe held by the blacksmith's tongs upon the anvil, and widened or narrowed at will, shows the chief advantage in which wrought iron excels. It can be forged and welded. When the carriagesmith joins two ends of a tire by heating them and then hammering them together he is welding and forging. The blacksmith can drag a hot iron rod from its red bed of charcoal and pull it out to a small diameter. In this same fashion all wrought iron can be pulled out into iron wire, or made into farming tools, rakes that gather in the harvest, or harrows that break the soil for the seed. This type of iron forms the huge links of the chains which hold vessels to their wharves or which raise great blocks of stone in granite quarries. It is the iron of the common nail. The production of wrought iron has been for many years one of the controlling interests of the iron trade, but at the present time the increase in production of steel, combined with its lowering in price, has brought about considerable replacement of the former by the latter. Wrought iron is the purest form of iron regularly manufactured. It contains the smallest quantity of carbon.

Steel. — Steel has more carbon in its composition than wrought iron and less than cast iron. It is this type of the metal which the world most uses to-day. Steel is more dependent on other elements for its structure than is wrought or cast iron, and it is often sold under a compound name indicating its composition. Such compound names are manganese steel, nickel steel, and the like.

The properties and tempering of steel. — Any one who has read the delightful story in "The Talisman" of the meeting between Saladin and Richard Cœur de Lion has a never-fading illustration of the wide-ranging possibilities of steel. Richard's mighty two-handed sword cut through the iron handle of the mace readily, but it could not cut the silken scarf. Saladin's sword of Damascus severed the floating tissue, but could make no impression on the iron. Both were good steel, but each had its own use. Steel possesses many good points. For instance, it is sufficiently fusible and malleable to be readily controlled in the furnace, while at the same time it is the hardest, strongest, and most serviceable of all the different varieties of iron. None of its properties provide such varying possibilities as the capacity which it possesses of being tempered: by tempering is meant the property by which steel varies in hardness according to the way in which the metal is cooled after being heated. The different action of Saladin's and Richard's swords came from the different rapidity with which each was cooled. Generally speaking, steel, when suddenly cooled, is brittle and hard. When slowly cooled, it is soft and tough. This is a very general statement, however, and we may say specifically that steel heated to about 450 degrees F., and cooled, gives the fine-cutting edge of razors. When similarly treated at about 500 degrees F. we get the metal used for shears, while 550 degrees F. gives the blue elastic steel of watch springs and of sword blades. The chilling commonly takes place in oil, and is only effective when iron is mixed with carbon. Perfectly pure iron is not tempered by sudden cooling. Wrought iron and cast iron exhibit this quality to some slight degree, but the percentage of carbon in steel seems to be the proper amount for the purpose.

Cast iron. — Cast iron is the metal as it comes from the furnace. It is the pig iron whose rough cylinders stand piled near foundries. Iron in this form contains the most carbon of all (and this is less than 5 per cent); as this type can be run into molds, it is used for all sorts of castings. The iron railings around the park; the iron pillars that hold up the elevated; the iron stove in the kitchen, — all these are made of cast iron. In this form the metal cannot be welded or forged.

Iron ores. — We said at the beginning of the chapter that iron ore could be reduced to a metal by heating the ore with carbon. That was the method used by early races in reducing the soft, pure iron ores. There are many ores from which iron can be extracted which require a more complicated method of separation. Iron is found in great plenty in generous deposits, scattered over practically the whole earth, and the red color which indicates the metal may appear in rock, clay, sand, or earth. The element exists in plants, in the blood, and, in its purest form, drops to the earth in meteorites coming from outer space. On earth pure metallic iron is seldom found, as the metal exists chiefly in combination with oxygen, carbon, or sulphur in the form of oxides, carbonates, or sulphides. The sulphides (iron pyrites, (FeS_2) etc.), useful as they have proved as a source of sulphur for sulphuric acid making, are seldom or never used as sources of metal supply. The carbonate ore is known as siderite, the oxide ores as hematite (Fe_2O_3) and magnetite (Fe_3O_4), while limonite, sometimes known as bog-iron ore, is not a pure oxide. Some of the properties of these iron compounds have been known for centuries. The power which magnetite, when found in the form of natural magnetic iron or lodestone, possesses of attracting to itself small scraps of metallic iron, as a horseshoe magnet attracts a needle, gave rise to many legends in olden times. The story of the mountain which drew all the nails from Sinbad's ship, as told in the Arabian Nights, was probably drawn from the narrator's memory of magnetite on such a mountain, and travelers even down to three hundred years ago steadfastly persisted in telling such fairy stories as veritable facts.

Occurrence of iron. — The United States is the great iron producer of all the world. Millions of tons are pouring yearly from our mighty blast furnaces, whose product is sent literally to every part of the globe. You will find American stoves and farming implements in the far north and at the equator. Most of the states possess more or less iron, but the Middle West, especially the lake region, and a small portion of the South contain the richest deposits. The placing of iron furnaces depends on more than iron deposits alone. Two other necessities for iron-making are coal and limestone. When the happy combination of all three things, iron, coal, and limestone, is found in any one place, that spot is likely to become a center of the trade. This is the reason why Birmingham, Alabama, is the foremost producer of all the cities of the South.

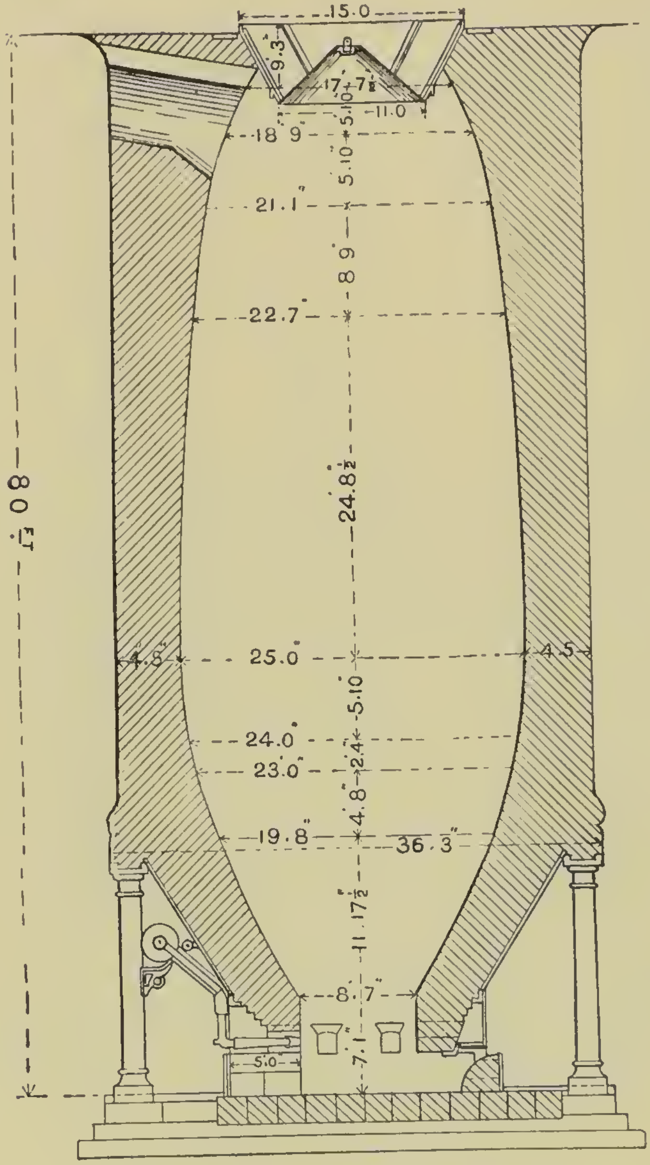
The smelting of iron. — Even to-day in the Pyrenees rich iron ore is smelted by means of carbon alone, in the same fashion that the armorers of Cæsar's legions used in smelting the iron for the short Roman swords. But most of our iron ore, while it will give up its oxygen to the reducing carbon, will not separate the metal from the clay with which it is mixed. The necessity for such division brings about the requirement of lime. This substance unites with the clay and forms a molten glass from which the metal can be separated. In three things, in the heating of the furnace, in the reducing action of the oxide formed from burning carbon (used here in the form of coal or coke), and in the formation of a glass from the clay of the ore by limestone, lie the main features of iron smelting. The glass or slag, as it is called, formed by limestone, silica, and some iron, is an essential part of the process. Ores which do not contain enough clay to give a good slag require the addition of silica as well as limestone to secure proper reduction.

The construction of iron furnaces. — The blast furnace of an iron plant towers into the air from fifty to one hundred feet, with a diameter at its largest point of from fifteen to twenty feet. Before smelting, iron ore is commonly crushed and roasted, a process which drives off water and carbon dioxide gas, makes the ores porous and more easily reduced. After roasting and crushing the broken mass is taken to the blast furnace. This furnace, like all other furnaces, is built to obtain the most complete combustion of the fuel used therein. It uses much fuel, and, in consequence, it needs much air. This

Blast Furnaces

air is supplied at a considerable pressure through pipes called tuyeres (pronounced tweers). The entering air is heated by the gases which are taken from the top of the furnace through an exit pipe and returned to the bottom of the furnace, there to give off heat. As one author has said: "The blast pipes correspond to the bellows used by a blacksmith and the exit pipe to the chimneys above his fire, except that gases escaping through chimneys are usually wasted." The blast furnace is made of iron-ribbed masonry or bricks, and is lined with fire-brick. When first started it has to be heated by means of fire, much as an ordinary furnace is heated, but after once starting it will run for a long time without cessation, and without other heat than that furnished by the carbon used in the charge. .Into the hot furnace is poured the charge, a mixture of the prepared ore, coke, and limestone, put on in layers like layer cake. The workmen add strata after strata of coal, ore, and flux (as the limestone is called, because it makes the clay a liquid or causes it to flow). The mixture goes in at the top and molten metallic iron flows out into a series of molds at the bottom. Between the beginning and end of the process many interesting changes occur.

The use of carbon in iron smelting. — Carbon is the substance that acts as the reducing agent of the iron, and the method by which this element takes the oxygen from the ore is very similar to the way in which coal acts in forming carbon monoxide in an open-grate fire. The air blast entering from the tuyere pipes meets the hot carbon, and the two form carbon monoxide due to the incomplete combustion of the carbon. It is not quite certain whether carbon dioxide is formed at this point in the furnace or not. In any case carbon monoxide is the gas which does the work, for as carbon dioxide rises in a furnace it passes over the red-hot upper fuel, and, in passing, is changed to carbon monoxide, the same blue flame gas that we see in an open fire. Carbon monoxide is always seeking for the oxygen needed to change it into carbon dioxide, and it finds that element in the oxide of the iron ore of the furnace. Iron exists in two forms, a ferric and a ferrous. The ferric oxide contains more oxygen than the ferrous. When the ore is composed chiefly of ferric oxide the carbon monoxide gas first reduces this compound to ferrous oxide and then reduces it still further to spongy metallic form. This spongy iron falls through the lighter masses about because of its greater density, and drops towards the bottom of the furnace taking up carbon as it falls. Finally it comes in a melted state to the bottom of the furnace, whence it may be drawn off.



A Blast Furnace in Section

The use of limestone in iron smelting. — The limestone of the charge meanwhile is having a set of reactions of its own. Limestone is largely calcium carbonate. We have already seen that calcium carbonate when heated breaks down into lime (calcium oxide) and carbon dioxide gas. The carbon dioxide which is formed goes through the same changes as that formed from the carbon of the fuel. The lime combines with the clay and the other silicates present in the ore, to form a glass — or slag as it is called. The formation of this slag removes the clay-like impurities from the iron, and, as it is of comparatively slight density, allows the heavier metal to sink through it. At the bottom of the furnace the molten iron is drawn off through holes in the base into molds (pig-iron molds). The lighter slag which accumulates above the iron overflows in a steady stream through an opening known as the slag hole.

Steel-making. The Bessemer process. — The usual process for making steel, the Bessemer (like its modification the Thomas-Gilchrist), depends for its success upon two factors: first, the removal of carbon from cast iron by burning the carbon out of the molten metal with a blast of air; second, the addition of a sufficient amount of carbon to the purified iron to form steel. The operation is carried on in a converter, a huge pitcher-shaped receptacle mounted on a carriage like a cannon carriage, on which the vessel can be swung to a horizontal or to a vertical position. The Bessemer process may be divided into four parts. In part one the converter is horizontal. It is filled with molten cast iron. In part two the converter is vertical. A strong blast of air is forced through the fused metal. This blast burns out carbon and other oxidizable impurities, leaving the iron free from carbon. In part three the converter is horizontal again. Enough cast iron (in the form which, containing manganese, is called spiegel iron) is added to bring the percentage of carbon in the steel to just the right amount. In part four the carbon of the spiegel iron is thoroughly diffused through the mass by heating. Steel is produced.

The Open-hearth process. — The Siemens-Martin, or Open-hearth process obtains the proper percentage of carbon for steel by mixing cast iron, having a heavy percentage of carbon, with wrought iron, having a low percentage of the same element. The mixture is then fused with an oxidizing flame. When the fused mass shows the right amount of carbon, ferromanganese is added. Steel is produced.

OPEN-HEARTH FURNACES

Chemical properties of iron. — The physical properties of iron and their effect upon its wide-spread use have already been mentioned. The chemical properties of the element are far less important. Iron generally forms ferrous and ferric salts as it combines with larger and smaller quantities of several elements. It readily forms oxides (such as the ores already considered), chlorides, sulphates, and sulphides. It forms nitrates with some difficulty. Both ferrous and ferric iron compounds are precipitated by such alkaline hydroxides as sodium hydroxide. This is a usual test for iron. The blood-red color which a sulphocyanide (a compound containing sulphur and cyanogen) imparts to ferric iron distinguishes between the two forms of this metal. The reactions for iron reduction are too complex in the main for use here. One may be given:

$$\underset{\text{Ferric oxide}}{Fe_2O_3} + \underset{\text{carbon monoxide}}{3CO} = \underset{\text{iron}}{2Fe} + \underset{\text{carbon dioxide}}{3CO_2}$$

Ferrous oxide = FeO. Ferric oxide = Fe_2O_3.

Ferroso-ferric oxide = Fe_3O_4. Ferrous chloride = $FeCl_2$. Ferrous sulphate = $FeSO_4$.

When these compounds are crystallized the first has four molecules of water of crystallization and may be written, $FeCl_2 \cdot 4H_2O$. The second has seven molecules of crystallization and may be written, $FeSO_4 \cdot 7H_2O$.

Ferric chloride (anhydrous) = $FeCl_3$. Ferric chloride (crystallized) = $FeCl_3 \cdot 12H_2O$. Ferric sulphate (anhydrous) = $Fe_2(SO_4)_3$.

COPPER

Copper the conductor of electricity. — Few movements of the last decades have been more sensational in their development than has been the conquering progress of electricity. Less than four decades ago the telephone was a toy but just conceived. Less than thirty years ago the first trolley car made its halting way along city streets. Less than two decades ago the first great plant destined to turn water-power into electrical power was begun and engineers began to change the energy of a current of water into the energy of a current of electricity. Now in the twentieth century electrical necessities and luxuries surround us on every side. To make machinery, and to send the current generated at central stations to the various points where it is needed, a substance is required through which the electrical current will flow with the least possible opposition or resistance. Such a conductor is copper.

Uses of copper in electrical work. — Copper is not the best conductor of electricity. First place belongs to silver, but copper is the best conductor that can be bought at a price that makes its general employment practicable. In consequence the wires that wind the field magnets and armatures of dynamos and motors, the bars and switches of the switchboards where men turn the electricity off and on, and the wires or cables that carry heavy electrical currents, are made from copper.

Properties of copper. — This great conductor of electricity is a red-colored metal in its normal state. It is hard and tough, easily bent, and readily drawn out into fine copper wire or thin copper foil. In moist air it soon tarnishes, turning to the iridescent green characteristic of copper carbonate. When heated to a high heat, copper changes to black copper oxide, turning the flame in which it is heated green. That green color with which copper in combination and in alloys tinges a flame, and the blue color which this element imparts to the solutions of most of its salts, are characteristic of the metal. Copper forms oxides, chlorides, sulphate, and sulphide. The formation of black copper sulphide with hydrogen sulphide is a common test for the metal dissolved in nitric acid. We can write this reaction as follows:

Copper nitrate + hydrogen sulphide = copper sulphide (ppt.) + hydrogen nitrate (nitric acid).

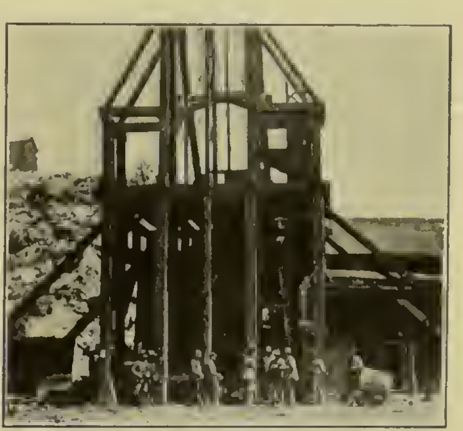
$$Cu(NO_3)_2 + H_2S = CuS + 2HNO_3.$$

General uses of copper. — Besides its use in electricity copper serves many purposes. It is the metal which constructs the round hot-water boiler in the kitchen. It is used for bolts and rivets on wooden yachts and ships, because tarnished copper will hold wood together where rusted iron fails. We use copper cents in the United States, copper pennies in Britain, copper centimes in France, and copper pfennigs in Germany. Films of copper are imprinted on plaster of Paris shells to make electrotypes, such as give us the illustrations for this book.

Alloys of copper. — The alloys of copper are next in importance to the metal itself. An alloy is a mixture or combination of metals which has been fused together. An alloy, therefore, is not a chemical compound in the ordinary sense of the word. Brass is probably the commonest alloy of copper. It sometimes contains almost seventy-five per cent of copper, the rest of its substance being zinc. Bronze such as we see in bronze statues contains tin as well as copper and zinc. Nickel coins, such as our five-cent piece, contain seventy-five per cent copper and twenty-five per cent nickel. The list of such alloys could be greatly enlarged.

Copper sulphate. — Copper sulphate ($CuSO_4$), the common "blue vitriol" of electric batteries, of calico printing, of dyeing, of insect-destroying sprays, and of copper-plating is the one copper compound we are most likely to meet. This salt of copper is formed when copper is treated with sulphuric acid.

Occurrence of copper. — Though by no means as widely distributed as iron, copper is a fairly abundant mineral. The deposits on the shores of the Great Lakes, in the mines of Montana, and in Arizona, furnish almost half of the whole supply of the world. There are rich deposits in Siberia, and mines are operated in Australia and in South America. There are also deposits in Britain, especially in Cornwall.



Copper. I

Hoisting works at a copper mine. Miners going down to work.

The reduction of copper from its ores. — The ores from which copper is reduced are its sulphides, oxides, and carbonates. In obtaining the metal the ore is first crushed and then washed to free it from impurities. The after treatment depends on the special ore which is to be treated. Carbonates and oxides are reduced by roasting with coke and fluxes. Sulphides are given a preliminary roasting in a reverberatory furnace. This burns off a large portion of the sulphides and changes them to oxides. The metal is then smelted in a blast furnace, by which process a "matte" is formed which contains copper mixed with other metals and sometimes with non-metals that may be present. A slag is formed which floats above the matte. The matte at this point is really an unrefined copper. The other bodies present beside the copper are then removed by a process known as "poling," which consists of stirring the molten mixture with long poles. The gases which come from the heating of the poles rise to the surface bearing with them the impurities of the matte. Properly done this gives about a 99 per cent pure copper.

The reactions for the different steps outlined above for the reduction of copper may be written as follows:

Simple reduction with coke — $Cu_2O + C = CO + 2Cu$.

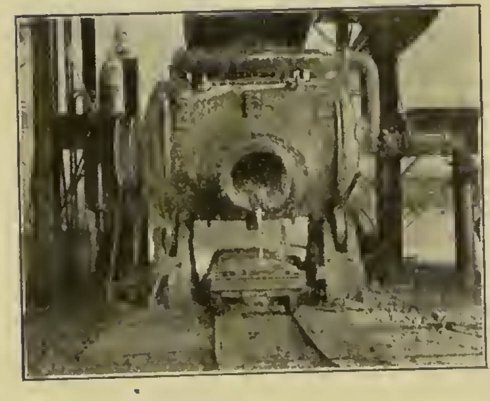
Fusing roasted ore — $Cu_2O + FeS = Cu_2S + FeO$.

$$2CuO + 2FeS = Cu_2S + 2FeO + S.$$

Roasting the sulphides with oxides in a reverberatory furnace — $2Cu_2O + Cu_2S = 6Cu + SO_2$.

At the same time iron is changed to oxide.

$$3Cu_2O + FeS = 6Cu + FeO + SO_2.$$



COPPER. II

Pouring molten copper from a converter into a mold.

The crude copper left may be still further purified by subjecting it to an electrolytic separation. By passing an electric current through two plates of copper, one of impure or crude copper, and one of pure copper, which are both dipped in a solution of a copper salt, a decided purification can be effected. The copper from the crude plate passes over and gathers as pure copper on the other plate. Copper made by this method is very pure and is called electrolytic copper. Among the copper ores are: Ruby ore = Cu_2O. Copper glance = Cu_2S. Azurite = $2CuCO_3$, $Cu(OH)_2$. Copper pyrites = Cu_2S, Fe_2S_3, or $CuFeS_2$.

CHAPTER XXXII

SODIUM [Symbol for sodium = Na]; POTASSIUM [for potassium = K]; CALCIUM [for calcium = Ca]; ALUMINUM [for aluminum = Al]; GOLD [for gold = Au]; and SILVER [for silver = Ag]

The earthy metals. Sodium, calcium, potassium, and aluminum. — Throughout the pages on which are spread the history of modern industrial chemistry are constant references to four metals of the soil, sodium, potassium, calcium, and aluminum. The salt of the sea, the potash of the woods, the marble of the quarries and the clay of the pits; all these are compounds of these four metals, for they are respectively the chloride of sodium, the carbonate of potassium, the carbonate of calcium, and the silicate of aluminum. The metals from which these substances are formed make up a group which possesses many singular characteristics.

Sodium

The color which sodium imparts to a flame. — When you drop sodium chloride into a flame, as when you spill a pinch of salt by accident into the stove, a yellow flame springs up. Its color shows the presence of sodium, and wherever a flame burns it is very likely to be tinged more or less by this metal, for particles of sodium exist scattered widely through the atmosphere which covers the sea and land. Sodium is not the only element that colors a flame. Every metal gives its own peculiar shade, when incandescent or glowing, and the flame test, as it is called, is a method of identifying metals by these characteristic colors. None are more vivid than the yellow of sodium and none more troublesome in the way of hiding other metals, for again and again sodium clouds the distinctive light of substances which we desire to test.

The electrolytic separation of sodium. — Metallic sodium does not occur in nature. It is prepared in the laboratory by the same method first pursued by the great chemist, Sir Humphrey Davy, whose researches in the first quarter of the nineteenth century formed such a brilliant prelude to the great symphony of discovery that followed. In 1807 Davy passed an electric current through fused sodium hydroxide and obtained metallic sodium. That was the first time the element had ever been isolated from its compounds. It remained a laboratory curiosity until, in very recent years, a modification of Davy's original scheme has become a commercial success. The general method of electrolytic separation, or separation by means of the electric current, has been discussed at some length before. It is sufficient for the present to say that all solutions that will transmit electricity can be separated into two parts, a positive and a negative. In the case of fused sodium hydroxide, sodium is the positive and hydroxyl the negative part. When the current passes through, the compound breaks down, the sodium goes to one side, the hydroxyl to the other, or, speaking more definitely, the sodium goes to the pole by which the current leaves and collects there. Hydrogen is set free at the same time as the sodium, and tends to keep the metal from oxidizing. The sodium may be removed as needed. The danger that the hydrogen formed may escape and explode makes this process altogether too dangerous for use in the ordinary laboratory.

The properties of sodium. — Sodium itself is a light silvery solid. The white color which we generally observe on the outside of sodium comes from the sodium oxide which forms whenever sodium is exposed to the influence of the air. We have already seen that sodium eagerly unites with water to form sodium hydroxide, setting hydrogen free, and that the rapidity with which it reacts with water is the chief difficulty experienced in keeping the metal. The moisture of the hand or even of the air is sufficient to start combining action, and the only way of properly retaining it in the metallic state is under oil. Sodium should not be touched with the naked hand. Few substances cause more painful burns. The metal is slightly lighter than water, floats in water, and melts a little below the boiling-point of water. By the aid of sodium some other metals such as magnesium are reduced, and certain sodium compounds formed from this element are of considerable value, particularly the peroxide of sodium which is used to bleach delicate fabrics.

Sodium compounds in fireworks. — One use of sodium belongs especially to joyous periods of celebration. When the crowd, gathered to watch a fireworks display, cries, "Oh!" and "Ah!" at the long trail of a sweeping rocket, the yellow spray of a sunburst, or the golden branches of a pot of flowers, few realize that the yellow fire which colors the black night is due to sodium. Sodium nitrate is used largely in making yellow fire or producing yellow flames.

Borax and boron. — If you ever dropped a bit of borax into the stove you saw the same yellow color rise that comes from salt. Borax is sodium biborate (formula, $Na_2B_4O_7 \cdot 10H_2O$) and is compounded of sodium, of oxygen, and of boron (symbol for boron, B), a non-metallic element which is much like silicon. Though boron is never found in elementary form, its compounds, boracic acid or boric acid (H_3BO_3) and borax, occur in enormous quantities in various places. Borax deposits are commonly far from the habitation of men, in wildernesses like the forbidden land of Thibet and the hot sands of Death Valley in California. Yet the wealth of borax in such deserts makes it profitable for men to undergo the privations

BORAX

Such teams as this bring borax through the desert.

necessary to obtain the powder and to overcome the gravest obstacles to transportation. Borax is used in glazing china, in enameling, and in making artificial jewels. It stands on the kitchen closet for use in cleansing. It cleanses metals from the oxides which would hinder soldering. It is used in welding metals and in dye-houses. It is also employed extensively as a preservative in canning and preserving foods. The grave question whether or not any chemical preservative of this type makes for the health of the people is still under investigation. We have scarcely sufficient data, as yet, to know definitely which are safe and which harmful.

POTASSIUM

The electrolytic separation of potassium. — Davy, who obtained metallic sodium by the action of the electric current, in similar fashion shortly after separated out metallic potassium, the metal nearest it in many properties. To-day this element is prepared from potassium hydroxide (KOH) in the same fashion that its sister metal is prepared from sodium hydroxide (NaOH).

Natural deposits of potassium compounds. — Very rich fields of natural deposits of potassium compounds exist near the old town of Magdeburg in Saxony, where lie the great Stassfurt deposits which are thought to be ancient beds of old primeval lakes. There in great layers, sometimes almost a mile in thickness, we find huge deposits of potassium salts left dry after the final evaporation of an inland sea. The metal itself is never found in nature outside of combination.

Properties and uses of potassium. — Potassium resembles sodium to a close degree. It is soft, light, and silver-white, and it floats upon the water. It is covered with a gray oxide under all ordinary circumstances, and is kept in oil because of its swift action on water. This metal releases hydrogen when brought in contact with water, and leaves potassium hydroxide behind. It combines in forms similar to those formed by sodium. In two characteristics, however, these differ. Potassium melts at 62½ degrees C. instead of 98 degrees C., and burns with a violet instead of with a yellow flame. The violet flame is the special test of metallic potassium. The use of potassium nitrate (KNO_3) in gunpowder, of potassium chlorate ($KClO_3$) in preparing oxygen, and of potassium carbonate (K_2CO_3) in making soap, will be recalled. Potassium hydroxide, like sodium hydroxide, is a strong base. It is used in manufacturing soft soap.

CALCIUM

Calcium, the metallic part of lime. — Common sight as freight cars laden with sulphuric acid are upon our railroads, cars filled with lime are, if anything, even commoner. Lime (calcium oxide) (CaO), the white powder, serves for buildings everywhere, and is a necessary article of supply to manufacturers of paper, of cloth, of iron, and of a thousand products more. The metal, calcium, never appears in elementary form in nature, but its oxide is a great commercial base used in those processes which demand a neutralizer or an alkali, in much the same way that sulphuric acid is used in those that require an acid. Slaked lime = calcium hydroxide = $Ca(OH)_2$. Unslaked lime = calcium oxide = CaO. Metallic calcium is golden yellow in color, is harder than lead, and malleable. It slowly unites with air and forms an oxide at ordinary temperatures. Swiftly uniting with air at high temperatures it forms an oxide and burns with a clear light. Calcium is heavier than water, but only between a fourth and a fifth as heavy as iron. Its characteristic flame test shows an orange.

The composition of the earth's crust. — Sodium comprises 2.5 per cent of the earth's crust, and the percentage of potassium is about the same; the weight of calcium in the earth is a per cent higher, 3.5; aluminum is present to the amount of 8 per cent. These four metals make up about 16.5 per cent of the whole. Oxygen in the form of oxides takes up about 50 per cent and 25 per cent is in the form of silicon. You can easily calculate from those figures that over 90 per cent of the earth's crust is made up of these six elements.

The preparation of calcium. — Calcium, like sodium and potassium, was prepared by Sir Humphrey Davy in the early part of the nineteenth century by aid of the electric current. It is now made by fusing calcium iodide with metallic sodium. The reaction for the process may be written thus:

Calcium iodide + sodium = calcium + sodium iodide.

$$CaI_2 + 2Na = 2NaI + Ca.$$

ALUMINUM

The preparation of aluminum. — Like sodium and potassium, aluminum is prepared by aid of the electric current. Never found free in nature, the metal is prepared from its oxides, which break down under the influence of electricity. Molten aluminum is thus formed in the presence of a fluoride and runs off into molds while the oxygen of the oxide unites with carbon to form carbon monoxide gas. In few cases are the possibilities of modern chemical research shown more clearly than in the manufacture of this metal. In two decades the discovery of inexpensive methods of production has lowered the price of aluminum to less than three per cent of its former cost.

The lightness and strength of aluminum. — Though heavier than calcium, aluminum is, nevertheless, one of the lightest metals in common use. It has a density of only 2.0. This light density combined with its strength, which is nearly as great as that of cast iron, makes the metal of great value where strength and lightness are essentials. This substance melts only at a high temperature.

Some uses of aluminum. — To recapitulate all the various uses which aluminum serves would take a catalogue as long as Homer's "Catalogue of the Ships," but a few places where we commonly meet the element may be mentioned. The frame of your camera; the cover on the preserve jars in the store-room closet; the bright, light saucepans which do not darken as the tin does; the silvery paint on the letter-box which so well withstood the weather, — all these are chiefly aluminum.

Gems which contain aluminum. — The old line, "The rose is red, the violet's blue," used sometimes to be given, "The ruby's red, the sapphire, blue." Probably the youthful poet did not realize that ruby and sapphire, topaz, amethyst, emerald, turquoise, and garnet all owed their living fire to this same element which is the element of clay, aluminum. All these jewels are compounds of aluminum crystallized under conditions of high temperature. Artificial rubies have been made from corundum and emery, those "hard clays," as they are sometimes called, found along the Atlantic coast of the United States. Corundum is but little softer than the diamond, and replaces it in much work where abrasive material is required to give a polish.

GOLD

Occurrence of gold and silver. — Gold and silver are metals of the mines rather than metals of the soil. Our immediate knowledge of them comes chiefly from their use in coinage and in jewelry. We are not likely to have much to do with their manufacture or their use in the arts. The desire for possession of these precious metals has been a ruling passion in every civilized or even semi-civilized nation since the dawn of history. Talents of gold and silver were used by the Greeks. Shekels of gold and silver were used by the Hebrews. The Spaniards entering Mexico found the golden ornaments among the Aztecs and were led on to exploration by their desire for the precious metal. It was the thirst for gold in 1849 that opened California and the Pacific Coast to the world, as the same thirst opened up the Klondike half a century later. Scattered as gold deposits are through the west of the United States, through Australia, South Africa, and Russia, there is never enough of it for the world's needs. Beyond all other metals it has been the goal of the treasure-seekers.

GOLD. I

Clearing up a gold claim in Alaska.

Washing gold. — This element appears in ores and in scales or nuggets in the ground. It is commonly found in alluvial soil, that is, in the earth deposits washed down by streams. Its separation depends upon the condition in which it is found in nature. In the early days of California and the West we find constant reference to "washing gold." That process depended upon the heavy density of the metal, which is over 19. A mixture of heavy gold and lighter sand, when shaken up with water, would have a tendency to separate. The lighter sand would rise to the top, the heavier gold would sink to the bottom. In time, if the washing was continued, all the gold would have collected on the bottom and could be separated. The essential principle of gold washing is separation made possible by the different densities present in a mixture. The heavier gold settles to the bottom through the lighter sand. The gold of alluvial soils in most cases originates in quartz veins whose disintegration has freed the precious metal.

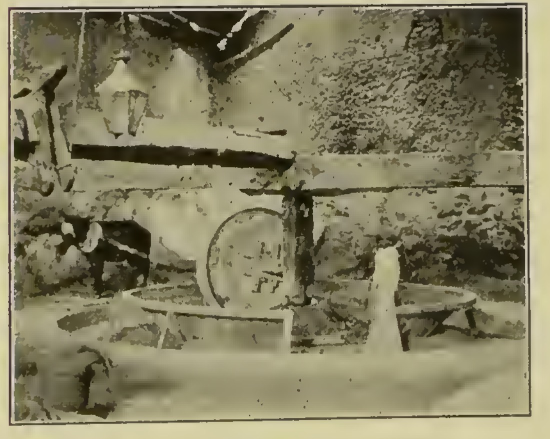
GOLD. II
Washing out gold with a gold pan.

The cyanide process for extracting gold. — When gold is found in quartz the ore receives a different treatment. It must be crushed and can then be treated by the cyanide process. When gold is exposed to the action of dilute potassium cyanide in contact with the air it forms a solution as a cyanide. (Solution of gold as a cyanide, $4Au + 8KCN + 2H_2O + O_2 = 4KOH + 4KAu(CN)_2$.) It is afterwards recovered in a metallic form by separation with the electric current or by precipitation of the gold upon zinc. In the cyanide reaction the water and oxygen from the air unite with four of the eight atoms of potassium to form four molecules of potassium hydroxide. The other four atoms of potassium join with four atoms of gold and eight radicals of cyanogen to form what is known as a double cyanide.

Measuring by carats. — Though a soft yellow metal, gold appears red in some lights. We meet it so commonly in the alloys of coins and rings that we naturally think of gold as being hard. As a matter of fact the hardness of such bodies is usually due to the copper present in them. The amount of gold in an alloy is determined by a standard known as the carat. Carats are expressed in parts per 24. Pure gold is 24 carats fine. A ring 18 carats fine would have 18 parts of gold and 6 of some other metal, such as copper or silver. A coin 22 carats fine would have 22 parts of gold and 2 of some other metal.

SILVER

Gold and silver standards. — Attempt after attempt has been made to fix silver in the place of gold as the world's standard of currency. No such effort has permanently succeeded. Economists have given various reasons for this, among them being the far larger production of silver than of gold. It is not impossible that the traditional rarity and precious quality of gold may have some weight in deciding the matter.



SILVER

Grinding silver ore in Mexico.

The permanence of silver. — The value of both the precious metals to jewelers and to makers of currency is due largely to the property which they possess of refusing to combine ordinarily with the oxygen of the air. Silver's tarnishing, as we have seen, is due to the action of sulphur.

Extraction of silver from its ores. — Silver is sometimes found in the metallic form, but far more commonly in combination as an ore. The usual way of extracting the metal is by the amalgamation process. Amalgams are those alloys in which mercury is present together with some other metal or metals. They are the simplest of the alloys to prepare, as they are formed by mixture of the constituents. When silver is extracted by one of the amalgamation methods the ores are powdered and silver chloride is formed by roasting or mixing the ore with salt (sodium chloride). This combination takes place readily. The silver chloride formed is then reduced to metallic silver by the action of iron. The equation which takes place might be written as follows in its simplest form:

Silver chloride + iron = silver + iron chloride.

$$2AgCl + Fe = 2Ag + FeCl_2.$$

At the end of this reaction the silver is left scattered through the whole mass. To separate it out mercury is added. Mercury, which will combine with gold or silver, but not with the other substances present, immediately forms an amalgam of mercury and silver. When this amalgam is heated the mercury, which boils at a far lower temperature than the silver, distills off and silver is left behind. Silver is a shining white metal which conducts heat and electricity better than any other metal in common use. Its power of conducting heat is seen whenever a silver spoon is left in hot water. Its power of conducting electricity may be perceived when the silver filling of a tooth is struck by a silver-plated fork. No property of silver is more striking than the changes which light produces in its compounds, but these have been mentioned in the chapter which treats of the influence of light upon chemical action. The test for silver when the metal is dissolved in nitric acid is the reverse of that already given for chlorine, the precipitation of silver chloride.

Silver nitrate + hydrogen chloride = silver chloride (ppt.) + hydrogen nitrate.

$$AgNO_3 + HCl = AgCl + HNO_3.$$

CHAPTER XXXIII

PAINTS, LEAD, TIN, ZINC, ANTIMONY, AND ARSENIC

Symbol for lead = Pb; for tin = Sn; for zinc = Zn; for antimony = Sb; for arsenic = As

Paints, useful for preservation and for beauty. — Probably the first sign of a community's desire for social advancement is its increasing use of soap. Perhaps the second is its free use of the paint brush. What a different impression is fixed upon a traveler's mind by a trim, well-painted farmhouse, church, or school, than is left by tumble-down, unpainted buildings, exposed to the destroying forces of the elements. The universal sway of the paint-brush comes first from desire for the preservation of wood or other building materials, and second from the desire for beauty or order. The paints which the brush spreads are metallic compounds held for general use in some convenient liquid medium. The pigments which cover the buildings of the earth come from mines where nature has stored metallic compounds.

A pigment and a dyestuff. — A pigment is quite a different matter from a dyestuff. A pigment is insoluble in water, and gives a mechanical coloring to the surface where it is applied. It acts either by adhesion (that is, by sticking to the surface) or by mixing (that is, by becoming mingled with the body on which it is placed). A dyestuff is soluble in water. It acts by forming a combination with some second substance. By this combination it deposits an insoluble colored substance within a textile fiber, producing a color by means of chemical action.

The application of a paint. Covering power. — Some medium which evaporates or hardens is necessary for the application of a pigment. White lead, for example, is applied mixed with linseed oil. The oil hardens and forms a coating with the lead. Shellac is dissolved in alcohol previous to use. It is applied in solution. After the application the alcohol evaporates and the shellac is left spread over the surface. A dyestuff soaks into the fiber of cloth by diffusion. We have already discussed this action of a dye in our consideration of dyeing with picric acid and with alizarin. A pigment must reflect the light and not transmit it, that is, a white paint must throw back a white light from the side of a house. The property which pigments possess of reflecting and not transmitting light is called covering power (a phrase we shall have occasion to use quite constantly by the way) and depends on the opacity or power which a pigment possesses of being impervious to the rays of light. To put it another way, covering power depends on lack of transparency. Dyestuffs, on the other hand, are sometimes required to both transmit and reflect light.

White lead. — Of all the metallic compounds used as pigments, white lead, basic lead carbonate ($2PbCO_3 . Pb(OH)_2$), is among the oldest and most generally used. This compound was known in the fourth century, and that splendid old chemist-alchemist, Basil Valentine, told of its adulteration in the fifteenth century. A double compound, it is both a carbonate and hydroxide of lead, and in recognition of this fact is called basic lead carbonate. Its wide use is due largely to four qualities: to the great ease with which it can be worked with oil, to the readiness with which it can be applied, to the completeness with which it covers, and to the power which it possesses of mixing with other pigments to form various tints.

The preparation of white lead. — Queer Dutch prints may even now be found which portray the old manufacture of white lead, a process which persists to-day after the passage of centuries. These prints show long sheds in which workmen in knee-breeches, buckled shoes, and flowing aprons are placing jars in long rows, which are composed of layer upon layer of queer rough earthenware pots, which look something like large, dull red flower-pots. The rows reach to the very roof. On one side a man is throwing thin sheets of metallic lead into the waiting receptacles, another is pouring on vinegar, and still a third is packing spent tan-bark all around the jars. In the end the jars are packed in tan-bark, something as eggs are packed in sawdust. The modern method is strikingly similar to the one just described, though manufacturers at the present time use gridiron-shaped pieces of lead instead of sheet-lead, and acetic acid instead of vinegar. For three months at a time the jars containing metallic lead and acetic acid are left exposed to air and water, under the action of the heat which comes from the fermenting tan-bark. All that time a whole cycle of reactions goes on between three things, the carbon dioxide, which is produced by the fermentation of the bark, the acetic acid present, and the metal. In the end white lead (basic lead carbonate) is left. By using a large number of small individual jars considerable quantities of lead can be treated at one time. In fact, from fifty to seventy tons are commonly converted at one treatment. Large quantities of white lead are prepared by the Lewis and Bartlett bag process, but the description of this method of manufacture is too complicated for insertion here.

White Lead Manufacture

The old style pot and the gridiron-shaped piece of lead placed therein.

Advantages of white lead. — Despite two disadvantages of white lead, the danger of lead poisoning, which is always present where lead is handled, and the further difficulty that its paint is readily blackened by sulphur compounds like hydrogen sulphide, no other pigment has been found which possesses so effective a covering power at so low a cost. Various newer methods of producing this compound, such as the passing of carbon dioxide into a solution of basic lead acetate, have been tried, but few have had the same success as the old method which is still in use to-day.

Zinc white. — Next in importance to white lead among the pigments is found zinc white, the oxide of zinc (ZnO), made in the simplest fashion by the direct burning of zinc in oxygen. This process is carried on in large furnaces where coal and zinc ore are burned together on a grate. The white zinc oxide powder produced is blown out through bag filters. This pigment is sometimes sold as "chinese white."

Advantages and disadvantages of zinc white. — The covering power of zinc white is much less than that of white lead. This pigment is not poisonous, and is not affected by sulphur fumes. Its power of resistance to the weather is poor. It requires much more oil than white lead, and, on the whole, is more expensive.

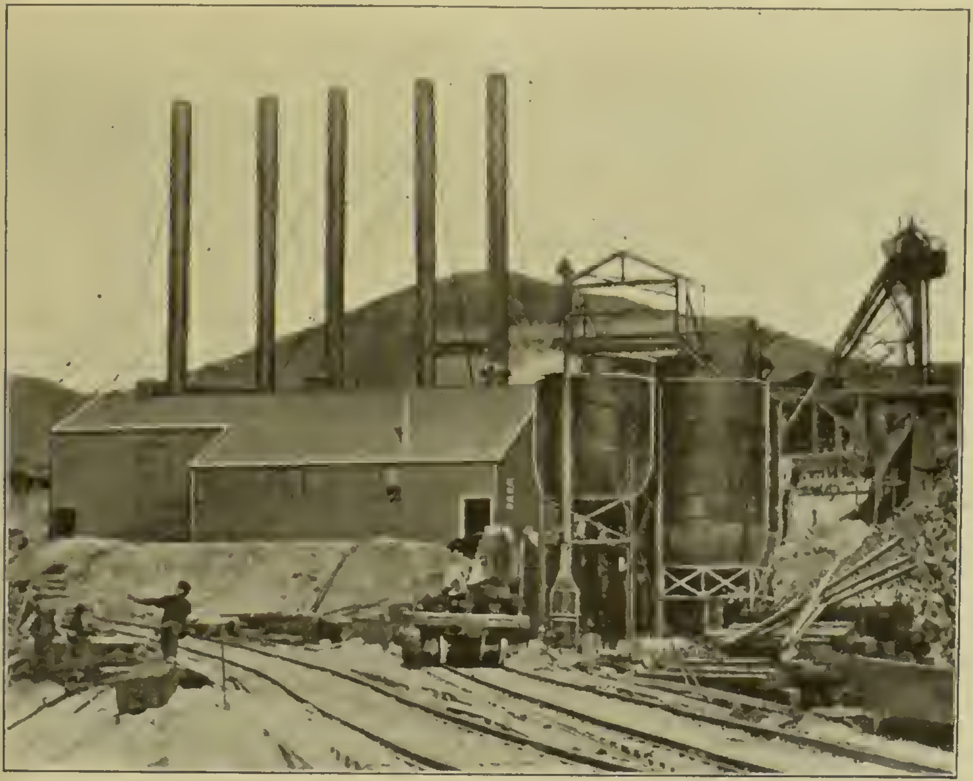
Blue, yellow, and red pigments. — Among the blue pigments are found salts of cobalt, copper and iron. The blue that showed in the invisible ink of cobalt chloride, the blue of copper salts, as well as the blue which comes from the combining of iron with the cyanogen radical, each and all take a share in producing the blue paints which color wall or curtain, theater scenery or motor car. (Prussian blue = ferric ferrocyanide = $4Fe(CN)_3 . 3Fe(CN)_2$). In the yellow pigments we find lead again taking a leading place in chrome yellow. Yellow ocher is an oxide of iron. Red lead, used in making red paints, is a lead oxide. (Red lead = minium (triplumbic tetroxide) = Pb_3O_4). Red antimony, used largely in rubber making, is a sulphide of antimony. The list might be largely extended.

LEAD

The softness and fusibility of lead. — If you have ever seen a plumber placing a lead pipe in your house, you have had an opportunity to see one of the qualities to which lead owes its widest use — its softness. From that quality comes the ease with which this metal can be bent, and the readiness with which it can be cut. To its fusibility, its property of melting at a comparatively low temperature, lead owes still other of its services. The manufacture of lead shot and lead bullets, of sheet lead and lead wire, depends on those two qualities of softness and fusibility. Lead shot has been made for many years by pouring molten lead through a sort of sieve at the top of a high tower. The fluid drops, which pass through, assume a spherical form as they descend and are chilled by the air as they drop into water below. When used for this purpose a small amount of arsenic is added to the lead. The use of lead in the preparation of strong acids is of great and increasing value to commercial chemistry. The metal is nearly insoluble in strong acids. Sulphuric acid up to a density of 1.58 has practically no effect on lead, and as a result the chambers in which this substance is prepared are lined with the metal.

The properties of lead. — Lead is bluish white in color in its normal state, a condition, by the way, in which we seldom see the metal, as it oxidizes swiftly in the air, forming a dark gray coating of lead oxide. It is so soft that it is scratched with a pin and leaves a mark behind when drawn over paper. Its density is 11.3. Because of its low melting-point, which is about 330 degrees C., lead can be readily reduced from certain of its ores. An attractive laboratory experiment shows the ease with which lead melts. If we take a bit of red lead (lead oxide or minium) on a hollowed piece of charcoal and submit it to a flame driven by a mouth blow-pipe, the carbon of the charcoal swiftly reduces the oxide to metallic form by the removal of oxygen to form carbon dioxide. In a few moments a small bead of metallic lead appears in the hollow, which begins to boil and bubble under the influence of the blow-pipe flame. When the flame is removed the lead solidifies and becomes covered with a film of oxide.

The reduction of lead from its ores. — Lead may be reduced from its commonest ore, galena (lead sulphide), by the reduction process. In this method the ore is roasted at a moderate temperature until, by the addition of oxygen, some of the sulphide is changed to oxide and sulphate. The temperature is then raised and the oxides mixed with sulphates in the furnace are heated out of contact with the air. By this means the sulphur present is forced to unite with the oxygen and the lead is obtained in the metallic state. In the more modern method the ore is first roasted in a reverberatory furnace to get rid of the sulphur and is then smelted in a blast furnace by a similar method to that described in the chapter on iron.



MINING
Ore bins and hoist.

The reduction process consists of two parts: (*a*) the roasting of the sulphide to sulphate and oxide.

$$PbS + 2O_2 = PbSO_4.$$
$$2PbS + 3O_2 = 2PbO + 2SO_2.$$

(*b*) The sulphate and oxide formed act on the remaining sulphide at a higher temperature.

$$PbSO_4 + PbS = 2Pb + 2SO_2.$$
$$2PbO + PbS = 3Pb + SO_2.$$

Lead poisoning from drinking water. — Lead salts taken internally are poisonous, and the chief danger of lead to the ordinary person lies in the fact that this metal, though insoluble in pure water, forms soluble compounds when water contains carbon dioxide, sulphates, carbonates, etc. When city reservoirs are charged with water containing these substances, soluble lead salts may be formed, which, taken internally, may cause dangerous illness. Waters which can dissolve lead should never be used after standing in the pipes. With any type of water it is a good plan to let a sufficient amount flow from the pipes to insure a reasonably fresh supply.

MINING

End view of a smelter for reducing metallic ores.

ZINC

The lightness and rigidity of zinc. — Zinc possesses one great advantage over lead — its lightness. Its greater rigidity is a second advantage. Zinc is only about one half as heavy as lead, taking it volume for volume (6.9 as against 11.5) and is especially serviceable where gutters are needed to carry water away from buildings or where water-pipes are required to carry overflowing water from roofs. Zinc is a very fusible metal. Only three metals, tin, lead, and bismuth, surpass it in this respect.

The reduction of zinc from its ores. — This element is seldom found in the metallic state, and is chiefly obtained from its sulphide ore, the mineral zinc-blende (ZnS). It is reduced from this ore by oxidizing both the zinc and the sulphur. The zinc forms zinc oxide, and the sulphur goes off as sulphur dioxide. The zinc oxide thus formed is mixed with carbon in the form of coke or coal and heated. The action of the burning reduces the zinc oxide to its metallic form, while carbon monoxide gas is formed from the carbon and oxygen. Preparation of metallic zinc from zinc-blende:

MINING

Side view of smelter.

$$ZnS + 3O = ZnO + SO_2.$$
$$ZnO + C = Zn + CO.$$

The properties of zinc. — Zinc is a brittle, bluish white metal. It crystallizes naturally and melts at 420 degrees. Its alloy with copper forms brass, as has already been mentioned. The addition of a small percentage of zinc in tin and antimony alloys has proved of value.

Tin

The "tin pan" and the "tin carpet tack." — The bright tin pan of the old pedler's cart and of the hardware store has really very little tin about it. A tin pan is generally a black sheet-iron pan, coated with tin, which has been made by dipping clean iron plates into hot liquid tin. The use of this bright metal as a covering to iron is due to its chemical properties. Since it does not oxidize in air and is not readily attacked by weak acids and alkalies, much cooking which would slowly destroy iron can be done in tin. Of late years the lowered price of enamel ware has caused a decline in the amount of tinware used in the kitchen, and the rows of shining tinware in which the housekeeper took pride have largely disappeared, driven out by the enamel lining, used for kettles and the like, which is a type of glass. The "tin" carpet tack is iron, coated with tin, and pins are made of brass covered with tin. Of the alloys of this metal, common solder, an alloy of tin and lead, is probably the most valuable. Solder is really a metallic cement used for joining two surfaces together. When a pan has broken at the edge, or an ice-chest lining leaks, a drop of solder cements the sides or fills the opening. It joins the ends of pipes with an insoluble and safe binding. Pewter, another alloy of tin, is made of three parts tin to one of lead. This really beautiful alloy makes rich lustrous ware. Cups for amateur prizes are commonly made of pewter. The use of tin in making bronze has already been considered in our study of copper.

The properties of tin. — Tin is a white metal soft enough to be cut with a knife. It melts at a low temperature, 228 degrees C., and crystallizes easily. It is slightly heavier than zinc, having a density in its ordinary condition of 7.2. Its crystalline form appears when dilute aqua regia is poured over a bright tin plate. Its reduction from its ores depends on three processes already discussed, roasting, washing, smelting.

ARSENIC AND ANTIMONY

Arsenic a metalloid. Where the danger of arsenical poisoning lies. — These elements stand on the border-line between metals and non-metals, where we find substances with some metallic and some non-metallic properties, elements sometimes called metalloids. Arsenic is metallic looking and conducts heat and electricity, but the compounds which it forms are more typical as a whole of the non-metals. It forms two oxides, arsenious and arsenic, and forms three acids from its higher oxide, the arsenic. The three acids formed from arsenic oxide have formulæ as follows:

Ortho-arsenic acid = H_3AsO_4.
Pyro-arsenic acid = $H_4As_2O_7$.
Metarsenic acid = $HAsO_3$.

This element forms compounds with all of the halogens, and like nitrogen it forms a direct compound with hydrogen, arsenic hydride known as arsine (AsH_3). Almost all the compounds of this element are dangerously poisonous, but arsine is perhaps more dangerous than any other. Whenever a compound of arsenic is in contact with decomposing organic matter it has the dangerous property of evolving arsine, caused by the growth of certain micro-organisms on organic matter which develop free hydrogen as they grow. The hydrogen thus formed enters into combination with any arsenic present and forms arsenic hydride. There lies the great danger of wall papers which contain arsenic. In a damp room invisible growths may gather on the walls and produce hydrogen there. If such a room is papered with arsenical wall-paper free hydrogen may form arsine in slow but steady quantities by combining with the arsenic of the paper. The poisonous effects of arsenic are cumulative and any one living in a room where arsine is being generated becomes gradually poisoned. Arsenic was not uncommonly found in wall-papers some years ago, but to-day laws have been quite generally passed forbidding the sale of arsenical wall-papers, and manufacturers have been forced to stop production. Arsenic is found in the earth in various mineral compounds, sometimes as sulphides, sometimes in native form or in iron pyrites which may contain arsenic in various proportions. To obtain the element some one of its minerals is heated in retort-like vessels. The arsenic becomes a vapor, passes off, and is condensed in cooled receptacles appearing at the end as a solid. When arsenical pyrites FeS_2, $FeAs_2$, is used as the ore from which to obtain arsenic, the reaction for its preparation goes on as follows:

$$FeS_2 \cdot FeAs_2 = 2As + 2FeS.$$

MINING
An old mine shaft.

Antimony. Its uses in its alloys. — Antimony's chief value rests on its property of forming alloys which expand in changing to a solid from the liquid state. That property of expansion gives even, sharp lines to castings, so that type metal made up of lead, antimony, and tin is of great value in making clear type. When the liquid molten alloy is poured into letter molds, each line of the letter comes out clearly, making a firm sharp impression. Antimony is a bright metal which does not oxidize at ordinary temperatures. When heated, it burns brilliantly, forming antimony trioxide. It forms a poisonous gas stibine which is antimony hydride (SbH_3). This compound is strikingly similar to arsenic hydride (arsine) in its poisonous qualities, in its general behavior, and in its constitution. Both arsine and stibine contain three atoms of hydrogen and one atom of the metals. Antimony occurs in nature chiefly as oxide and sulphide. It can be reduced from its sulphide by heating it in plumbago crucibles with scrap iron. Under these circumstances the sulphur present combines with the iron and frees antimony. The reaction is as follows:

Antimony sulphide + iron = antimony + iron sulphide.

$$Sb_2S_3 + 3Fe = 2Sb + 3FeS.$$

White antimony = Sb_2O_3.

CLOSE

From our knowledge of the almost infinitely tiny ion, from the gaseous mantle which covers the earth, from great factories, wide-spreading fields, and restless seas, the pages of this book have been gathered together.

To-day the past shows a great mass of ordered chemical knowledge which for decades has been transforming the world. The great material future must be developed largely through the sister sciences of physics, chemistry, and biology. To these we must look for the greatest possibilities of those achievements which shall bring health, happiness, and material wealth to mankind.

In our own field of chemical research three lines give especial promise. Catalysis and electrolytic dissociation on the side of applied chemistry, and the revelations of radioactive matter on the side of theoretical chemistry. Of the first two we have spoken briefly. The third involves more complex problems, more wide-spreading theories than can properly be discussed here. Ten years more, even five years more, may bring results which may sufficiently order our knowledge of radium, of its emanation, of helium, and of the interrelation between the elements, as to enable us to enter a discussion of radioactive matter in the elementary text-book. To-day it is not feasible.

The work of Ramsay and Cameron, of Rutherford and of Soddy, of Madam Curie and of Crookes, is epoch-making in its greatness, following as it does the labors of those brilliant pioneers, Becquerel and Monsieur Curie. Great as that work is, it is yet in the making, and discussion of the results obtained and of the possibilities of the future must be postponed.

It has been said that the writing of history never stops while the printing-presses of newspaper and magazine rattle and clash. It is quite as true that the writing of chemistry never stops while the press records the progress of chemical knowledge produced by new and adventurous explorations into the unknown lands of science.

Once more let us say, that if this book serves as an incentive to further scientific study and as a guide to the wealth of industrial knowledge which is pouring out upon us daily, it will have achieved one of its specific aims. No field of human effort yields richer harvests than those new worlds which science is ever opening to our view and for our use.

CHAPTER XXXIV

CHEMICAL ARITHMETIC

Chapter XXXIV. on chemical arithmetic may be distributed over the course. Density and specific gravity may be considered with Chapter II. and Chapter III. Percentage of water of crystallization may be considered with Chapter VI. on water. Percentage of composition may be considered with any of the five chapters following Chapter XV. on law. Calculation of gram weights having given molecular weight should follow the percentage of composition problems directly. The work in thermometry should follow Chapter XXXIII. on heat and light, and the discussion of Charles's and Boyle's law should follow Chapter XXI. on the gas laws. Determination of simplest formulæ may well end the work in chemical arithmetic. As no previous training in science is assumed some of the arithmetic of physics has been introduced into this chapter. The only knowledge assumed is some training in decimals. As many pupils are deficient in this subject it may be well to preface the study of chemical arithmetic with a brief review of decimal fractions.

The definition of density is as follows: — Density is the amount of matter per unit of volume. Two factors must be known to find density (1) amount of matter, (2) the unit or units of volume.

In determining density we use the balance to ascertain the number of grams or of pounds in the body since weight measures amount of matter. To calculate volume we use one of the two common units of volume, the cubic centimeter or the cubic foot.

A cube which has a volume of 4 cubic centimeters and a weight of 8 grams would have a total amount of matter of 8 grams. The total number of units of volume would be 4 cc. There would then be 2 grams to every cubic centimeter. This appears from the following diagram in which each space represents 1 cc.

2 gm.	2 gm.
2 gm.	2 gm.

8 gm. = weight. 4 cu. cm. = volume

$$\text{Density} = \frac{8 \text{ gm.}}{4 \text{ cu. cm.}} = \frac{2 \text{ gm.}}{1 \text{ cu. cm.}}$$

Since dividing the weight by the volume gives the weight for each unit of volume, we may write the definition of density as a formula:

$$\text{Density} = \frac{\text{Weight}}{\text{Volume}} \text{ or } D = \frac{W}{V}.$$

$$\text{If we use the metric system: Density} = \frac{\text{Number of grams}}{\text{Number of cu. cm.}}.$$

$$\text{If we use the English system: Density} = \frac{\text{Number of pounds}}{\text{Number of cu. ft.}}.$$

As a typical example, suppose we take a cube of a metal which weighs 84 grams, has a length of 4 cm., a breadth of 2 cm, and a thickness of 3 cm. What is the density of this cube?

Density $= x$. Weight $= 84$ grams.

$$\text{Volume} = \frac{\text{Length}}{4\text{ cm.}} \times \frac{\text{Breadth}}{2\text{ cm.}} \times \frac{\text{Thickness}}{3\text{ cm.}} = 24\text{ cu. cm. Density} = \frac{\text{Weight}}{\text{Volume}}.$$

Substitute for the general terms weight and volume the values given above and $x = \frac{84}{24}$ density $= \frac{3.5}{1}$.

We should probably write this answer as 3.5, omitting the figure 1 which stands in the second part of the ratio, but we should always understand by such a numeral that there are 3.5 grams in every cubic cm. of the substance. In similar fashion we can figure the density of any body whose weight and volume are known. A body containing 20 cu. cm. and weighing 80 grams is seen at once to have 4 grams to each cubic centimeter and therefore has a density of 4.

This formula serves to give us weight and volume as well as density. For example: if the density of a body is 6 and its volume 20, we find weight by substituting these values in the formula.

$$\text{Density} = \frac{\text{Weight}}{\text{Volume}}. \quad 6 = \frac{x}{20}. \quad x = 6 \times 20 = 120 \text{ grams.}$$

If we know the density and the weight, we can figure the volume. Suppose we take a body whose density is 4 and whose weight is 20.

$$\text{Density} = \frac{\text{Weight}}{\text{Volume}}. \quad 4 = \frac{20}{x}. \quad 4x = 20. \quad x = 5.$$

PROBLEMS ON DENSITY

1. A piece of wood is 24 cm. long by 11.5 broad by 8.3 thick, its weight is 1380 grams. Find its density.

2. A liquid has a volume of 2.56 liters, it weighs 3120 grams. Find its density.

3. A cube of an alloy is 82 cm. long by 26 broad by 10.6 cm. thick, its density is 4.3. Find its weight.

4. A piece of china weighs 187 grams, its density is 2.7. What is its volume?

5. A liquid has a volume of 4.75 liters, its density is .7. Find its weight.

6. A cube of wood weighs 146 grams, it has a volume of 432 cu. cm. Find its density?

7. A liquid has a volume of 392 cu. cm., it weighs 287.4 grams. What is its density?

8. A cube of a certain alloy is 2.2 cm. long, 1.7 cm. wide and .3 cm. thick, it weighs 19.60 grams. Find its density.

9. A china vase weighs 44 grams and has a density of 2.35. Find its volume.

10. A liquid has a volume of 11.4 cu. cm., its density is .72. Find its weight.

Specific Gravity

Specific gravity is the ratio which exists between the weight of any body and the weight of a like body of water or, stating it in another fashion, it is a comparison between the weight of any fixed volume of a liquid or solid compared with the weight of a like volume of water.

Writing this as a formula: $\text{Sp. Gr.} = \frac{\text{Weight of body}}{\text{Weight of like body of water}}$.

Numerically, specific gravity calculated by the metric system is no different from density because the weight of 1 cu. cm. of water = 1 gram, so that the volume of any body = the weight of a like body of water. If the volume of a body, for instance, is 8 cu. cm. the volume of a like body of water (that is, of a body of water of equal size) would be 8 cu. cm. and its weight would be 8 grams. If the weight of the body were 16 grams, its density would be $\frac{16 \text{ grams}}{8 \text{ cu. cm.}} = \frac{2}{1}$.

Its specific gravity = $\frac{16 \text{ grams, weight of body}}{8 \text{ grams, weight of like body of water}}$. Since 8 is the denominator of both of these fractions, the density = 2, the specific gravity = 2. In the first case, however, we have been dividing by 8 cu. cm., in the second by 8 grams. The first case states the comparison which exists between the weight and volume of a single substance. The second case states the comparison which exists between the weights of like volumes of two different bodies.

If a body weighed 1800 grams and had a volume of 900 cc., we could find its specific gravity by the formula given above since a like volume of water to that given (900 cc.) would weigh 900 grams. The formula then becomes Sp. gr. $= \frac{1800}{900} = \frac{2}{1}$. The body weighs twice as much as the same amount of water.

In the English system specific gravity and density are two different quantities numerically. The weight of one cubic foot of water is 62.4 pounds. If a body weighed 124.8 pounds and had a volume of 1 cu. ft., its density $= \frac{\text{Weight}}{\text{Volume}} = \frac{124.8}{1}$.

Its specific gravity

$$\frac{\text{Weight of a cu. ft. of the substance}}{\text{Weight of a cu. ft. of water}} = \frac{124.8}{62.4} = \frac{2}{1}.$$

For example, a body is 4 ft. long, 2 ft. wide, and 1 ft. thick, it weighs 720 pounds, what is its specific gravity?

The specific gravity $= x$. Weight $= 720$. The weight of a like body of water $8 \times 62.4 = 499.2$. $x = \frac{720}{499.2}$. Sp. Gr. $= \frac{1.4}{1}$.

A method for finding the weight of any body having given the specific gravity and the volume follows. If the specific gravity of a body is 4 and the volume 4 cu. ft., we can be sure that the body is four times as heavy as water. 4 cu. ft. of water would weigh 62.4 lbs. × 4 = 249.6. This body weighs four times as much as water, so its weight would be 249.6 × 4 = 998.4.

PROBLEMS ON SPECIFIC GRAVITY

1. Find the specific gravity of a body whose volume is 2400 cc. and whose weight is 920 grams.
2. Find the specific gravity of a body whose volume is 3 cu. ft. and whose weight is 1246 pounds.
3. Determine the unknown quantity in the following examples:

Volume	Weight	Specific Gravity
x	24 grams	3
x	129 pounds	5
28 cc.	180 grams	x
1.5 cu. ft.	380 pounds	x
399 cc.	x	4.2
920 cu. ft.	x	3

4. A piece of iron with a volume of $\frac{1}{3}$ of a cubic foot weighs 164 lbs. Find the specific gravity.

5. What is the specific gravity of a body, volume 2 cu. ft. weight 248 lbs.? (One cu. ft. of water weighs 62.4 lbs.)

6. If 1 cu. ft. of water weighs 62.4 lbs. and 7 cu. ft. of alloy weigh 1400 lbs., what is the specific gravity of the lead?

Percentage of Water of Crystallization

Many chemicals take up more or less water, from the air, or from solutions, which enters into their structure, in the form of water of crystallization. Many of these chemicals are sold on the basis of their dry weight, that is, of their weight when water has been driven off. To calculate this dry weight it is necessary to know the percentage of water of crystallization present. For example, if a sample of barium chloride weighed 244 grams, part of that weight would be dry barium chloride and part of it would be water of crystallization. If we heat the 244 grams until all the water is driven off, the dry powder left will weigh 208 grams. Subtracting, 244 − 208 = 36 grams. This 36 grams is water. Under these circumstances it is water of crystallization. The 208 grams which were left after heating is dry barium chloride. To find the percentage of water of crystallization we divide the weight of water of crystallization, 36 grams, by the total weight, 244. 36 ÷ 244 = .147. .147 × 100 = 14.7 per cent water of crystallization.

Note.—We multiply by 100 because percentage must be stated in parts per 100. If we divide the partial weight by the whole weight and do not multiply by 100 we get a decimal.

The percentage of dry barium chloride would be

$$\frac{\text{the weight of dry barium chloride}}{\text{total weight}} = \frac{208}{244} \times 100$$

or approximately 85.5 per cent.

Model Problem on Water of Crystallization

17 grams of copper sulphate lost 3.5 grams on heating. What was the per cent of water crystallization lost?

$$\frac{3.5 \text{ grams} = \text{amount lost}}{17 \text{ grams} = \text{whole amount}} \times 100 = \text{per cent lost.}$$

```
17)3.5(.205 × 100 = 20.5 per cent.
   3 4
   ---
     100
      85
      --
      15
```

20.5 = per cent of water of crystallization lost.

1. Calculate the per cent of water of crystallization in a sample of barium chloride, 1.5 grams of which lost on heating 0.22 grams.

2. A student found that 2 grams of barium chloride lost on heating 0.295 grams. What per cent of the sample was water of crystallization?

3. A sample of crystallized copper sulphate weighing 3.6 grams lost 1.4 grams on heating 10 minutes and .7 grams more on heating 20 minutes. What per cent was lost in each case and what was the total per cent of crystallization (water of crystallization) in the salt?

4. 27.8 grams of ferrous sulphate contain 712.6 grams of water. Calculate the per cent of water of crystallization in the compound.

5. In iron alum practically 43 parts by weight out of 81 parts are water. Calculate the amount of water that 28 grams of this substance would lose when thoroughly dried.

6. Chrome alum has 51.4 parts by weight of dry salt out of 94.6 parts by weight of crystallized salt. How much water would 28 grams of this substance which contains chromium instead of iron lose when thoroughly dried?

Percentage of Composition

By a similar process to that last stated we can obtain the percentage of composition of any chemical compound. To obtain the total percentage of each element present in a compound the total weight and the weight of each part must be known. When we know those factors we can divide the weight of each part by the total weight and then multiply by a hundred to get the per cent.

NOTE.—This calculation of percentage is the same as that just performed to obtain percentage of water of crystallization.

Model example: We have 98 grams of sulphuric acid, 2 of these are hydrogen, 32 sulphur, 64 oxygen. Find the percentage of composition: $\frac{2}{98}$ of the whole will be hydrogen, $\frac{32}{98}$ sulphur, and $\frac{64}{98}$ oxygen.

$$\frac{2}{98} = .0204 \times 100 = 2.04 \text{ per cent hydrogen.}$$

$$\frac{32}{98} = .3265 \times 100 = 32.65 \text{ per cent sulphur.}$$

$$\frac{64}{98} = .6530 \times 100 = \begin{array}{l} \underline{65.30 \text{ per cent oxygen.}} \\ 99.99 \text{ per cent} = \text{total per cent.} \end{array}$$

NOTE.—In calculating these percentages an answer like this is quite near enough. We seldom or never obtain 100 per cent in practice.

A second use of such problems may be found in determining the amount of an element present in any quantity of a substance. For example: 18 grams of water is composed of 2 grams of hydrogen and 16 grams of oxygen. How much oxygen will be present in 194 grams of water?

$\frac{16}{18}$ of the whole quantity will be oxygen. $\frac{16}{18} = .8888 \times 100 =$ 88.88 per cent. 194 grams × 88.88 per cent = 172.43 grams.

It is evident from the examples shown above that we can find either the percentage of any element present in a compound or the total percentage of all the elements present. We might write formulæ for such examples as follows:

$$\text{Percentage of composition of any part} = \frac{\text{Weight of part}}{\text{Total weight}} \times 100.$$

Percentage of composition of whole = sum of per cents of all parts.

Second Model Problem on Percentage of Composition

How much calcium in 200 grams of calcium carbonate ($CaCO_3$)?

$$Ca = 40 \cdot C = 12 \cdot O = 16$$

$$\begin{array}{rl} Ca & = 40 \\ C & = 12 \\ 3O & = \underline{48} \\ & \ 100\ CaCO_3 \end{array}$$

$$40 \times \frac{200}{100} = \frac{8000}{100} = 80 \text{ grams.}$$

$\frac{40}{100}$ of any amount of $CaCO_3$ is Ca.

80 grams of 200 grams $CaCO_3$ is Ca.

There are 80 grams of calcium in 200 grams of calcium carbonate.

1. 63 grams of nitric acid is composed of one gram hydrogen, 14 grams nitrogen, and 48 oxygen. Find its percentage of composition.

2. 46 grams of alcohol is composed of 24 grams carbon, 6 grams hydrogen, 16 grams oxygen. Find its percentage of composition.

3. 58.5 grams of salt is composed of 23 grams salt, 35.5 grams chlorine. Find its percentage of composition.

4. 216 grams of oxide of mercury is composed of 200 grams mercury, 16 grams oxygen. What is the percentage of mercury in the compound?

5. 124 grams of a wet chemical is throughly heated and loses 28 grams. What per cent of the total weight is water?

Note.—The word parts in the following examples means parts by weight.

6. Plaster of Paris is composed of 40 parts calcium, 32 parts sulphur, 64 parts oxygen. Find the amount of calcium in 195 grams of plaster of Paris.

7. Anhydrous copper sulphate is composed of 63 parts copper, 32 parts sulphur, 64 parts oxygen. Find the per cent of oxygen in 108 grams of this compound.

8. What per cent of oxygen is in 15 grams of potassium chlorate if 48 parts out of 132.5 are oxygen?

9. Carbon dioxide is composed of 12 parts carbon, 32 parts oxygen. What per cent of each is 236 grams of carbon dioxide?

10. Calcium carbonate is composed of 40 parts calcium, 12 parts carbon, 48 parts oxygen. Find the per cent of each in 632 grams.

Calculations of Gram Weights Having Given Molecular Weights

One of the most valuable of the formulæ of elementary chemistry is based upon the law of constant proportions by weight. Wherever we find the molecule of salt, NaCl, it is always made up of an atom of sodium which weighs 23 microcriths and an atom of chlorine which weighs 35.5 microcriths. Wherever we find a molecule of mercuric oxide, HgO, it is made up of an atom of mercury weighing 200 microcriths and an atom of oxygen which weighs 16 microcriths. The total weight of the molecule HgO is 216 microcriths.

According to the law of constant proportions, substances can decompose only in the ratio of their atomic weights. If we have 216 grams of HgO we shall have 200 of mercury and 16 of oxygen when the compound is decomposed by heat. We shall have 8 of oxygen and 100 of mercury when one half of that quantity is decomposed by heat. However much mercury oxide we may have it will always break down so that $\frac{16}{216}$ of its total weight is oxygen and $\frac{200}{216}$ is mercury. In short its parts will always appear in ratios which bear a simple relation to its atomic weights. When sodium chloride is decomposed by the electric current $\frac{23}{58.5}$ of any quantity whatsoever appears as sodium and $\frac{35.5}{58.5}$ of any quantity whatsoever appears as chlorine. All compounds are invariably constant in their composition. Like action takes place with every chemical compound placed under similar circumstances wherever that compound is found. In consequence we may make a proportion as follows:

The molecular weight of the whole of any body: molecular weight of any part of that body: : gram weight of the whole: gram weight of a part.

To take a specific example, suppose we have 54 grams of mercuric oxide, HgO, and we wish to know how much oxygen it contains. We can determine the amount desired by inserting the known molecular weights and the known gram weight in their proper places in the proportion. The proportion would then read as follows:

The molecular weight of the whole: molecular weight of the part: : gram weight of the whole: gram weight of the part.

$$\begin{aligned} HgO &= 200 + 16 : O = 16 : : 54 \text{ grams} : x. \\ & \quad 216 : \quad 16 : : 54 \quad : x. \\ 216\,x &= 54 \times 16. \\ x &= 4. \end{aligned}$$

There will be 4 grams of oxygen in every 54 grams of HgO.

NOTE.—The typical proportion runs as follows:

$$a : b :: c : d.$$

a and d = the extremes.
b and c = the means.

Since $ad = cb$ the products of the means is equal to the product of the extremes.

This law just considered will hold just as good with pounds as with grams. It uses atomic weights as well as molecular weights, but for convenience it is well to remember the general form given above. The atomic weight of all the elements will be found in a table on the cover page.

In calculating the molecular weights of somewhat more complex compounds such as crystallized barium chloride, $BaCl_2 \cdot 2H_2O$ the computation is performed by adding together the atomic weights of each of the separate elements. $(Ba = 137)$ $(Cl_2 = 2 \times 35.5 = 71)$ $(2H_2 = 4 \times 1 = 4)$ $(2O = 2 \times 16 = 32)$ $137 + 71 + 4 + 32 = 244$. Total = 244.

NOTE.—Use the approximate figures in the table on the cover page wherever atomic weights are required in the succeeding problems. Write the formula in symbols as the first step in every case.

PROBLEMS

1. How much hydrogen in 17 grams of sulphuric acid?
2. How much oxygen in 22.5 grams of potassium chlorate?
3. How much sulphur in 22 grams of hydrogen sulphide?
4. How much chlorine in 87 grams of hydrochloric acid?
5. How much nitrogen in 84 grams of ammonia?
6. How much carbon dioxide in 62 grams of calcium carbonate?
7. How much iodine in 17 grams of potassium iodide?
8. How much carbon in 24 grams of methane?
9. How much iron in 248 grams of ferrous sulphate?
10. How much silver in 106 grams of silver nitrate?

This formula holds good not only in the composition of individual compounds but also in any type of reaction whether such reactions are synthetical, analytical, or metathetical.

Suppose, for example, we wish to know how much silver chloride we can make from 106 grams of silver nitrate. In order to make silver chloride we should naturally add sodium chloride, $NaCl$, to silver nitrate, $AgNO_3$. The reaction would be as follows:

Silver nitrate + sodium chloride = silver chloride + sodium nitrate.

$$AgNO_3 + NaCl = AgCl + NaNO_3.$$

The atoms in this reaction can only transfer themselves bodily and can only react in their atomic weights. Therefore we can write a proportion practically the same as that which we wrote before.

The molecular weight of one part of the reaction: molecular weight of any other part of the reaction: : gram weight of the first part: gram weight of the second. If we write the reaction above, inserting the atomic weights, it would read as follows:

$AgNO_3$	$NaCl$	$AgCl$	$NaNO_3$
$Ag = 108$			$Na = 23$
$N = 14$	$Na = 23$	$Ag = 108$	$N = 14$
$3O = 48$	$Cl = 35.5$	$Cl = 35.5$	$3O = 48$
$AgNO_3 = 170.$	$NaCl = 58.5.$	$AgCl = 143.5.$	$NaNO_3 = 85.$

Only two of these molecular weights concern us here: the weight of the silver nitrate and the weight of the silver chloride. These we will substitute in our proportion, which now reads: 170 (the molecular weight of $AgNO_3$): 143.5 (the molecular weight of $AgCl$): : 106 (the gram weight of $AgNO_3$): x (the gram weight of silver chloride).

$$170 : 143.5 : : 106 : x.$$
$$170x = 143.5 \times 106.$$
$$170x = 15211.$$
$$x = 89.4 \text{ grams of } AgCl.$$

Problems

1. How much zinc is required to form 63 grams of zinc sulphate?

Equation: $Zn + H_2SO_4 = ZnSO_4 + H_2.$

2. How much hydrochloric acid is required to form 28 grams of sodium chloride?

Equation: $HCl + NaOH = NaCl + H_2O$.

3. How much sulphuric acid will form 54 grams of potassium sulphate?

Equation: $H_2SO_4 + 2KOH = K_2SO_4 + 2H_2O$.

4. How much carbon dioxide can be made from 87 grams of calcium carbonate?

Equation: $CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$.

5. How much ammonium hydroxide will be necessary to make 18 grams of ammonium chloride?

Equation: $NH_4OH + HCl = NH_4Cl + H_2O$.

6. How much sulphuric acid will make 127 grams of nitric acid?

Equation: $H_2SO_4 + 2NaNO_3 = 2HNO_3 + Na_2SO_4$.

7. How much hydrogen sulphide will form 86 grams of copper sulphide?

Equation: $H_2S + Cu\,(NO_3)_2 = CuS + 2H\,(NO_3)$.

8. How much carbon dioxide and water can be obtained from 28 grams of methane?

Equation: CH_4 (completely oxidized) $= CO_2 + 2H_2O$.

THERMOMETRY

The commonest calculations of thermometry are the changes to Centigrade degrees of temperatures written in Fahrenheit degrees and vice versa. It is sometimes advisable to change Centigrade degrees to absolute degrees. There are 100 degrees between the freezing- and boiling-point of water on the Centigrade thermometer, and there are 180 degrees between the freezing- and boiling-point of the Fahrenheit. In consequence each Centigrade degree is $\frac{9}{5}$ of a Fahrenheit degree. Each Fahrenheit degree is $\frac{5}{9}$ of a Centigrade degree. That is to say, since there are 180 Fahrenheit degrees to 100 Centigrade degrees, each single Centigrade degree must be 1.8 or $\frac{9}{5}$ as great as the Fahrenheit. Since there are 100 Centigrade degrees to 180 Fahrenheit degrees, each Fahrenheit degree must be $\frac{1}{1.8}$ of a Centigrade degree, or it must be $\frac{5}{9}$ of a Centigrade degree. These facts are made more evident by the Centigrade scale which follows.

Since all calculations must be reckoned from the same starting-point, any Fahrenheit degree must have 32 degrees subtracted before its corresponding Centigrade degree can be obtained. This is necessary because 32° F. = 0° C. An absolute degree is equal to a Centigrade degree, but its zero is taken 273 degrees below zero Centigrade, and begins at absolute zero. To change Centigrade degrees to absolute degrees we may add 273 degrees, since 0° C. = 273 absolute.

I. Fahrenheit — Boiling 212 — Freezing −32
II. Centigrade — Boiling 100 — Freezing 0
III. Absolute — Boiling 373 — Freezing 273

I.	II.	III.	
212°	100°	373°	Boiling-point
32	0	273	Freezing-point
		0	Absolute zero

In going from the freezing-point to the boiling-point of water the Centigrade scale has 100°. There are a hundred steps on this ladder of temperature. A Fahrenheit thermometer has its freezing-point at 32°; its boiling-point is at 212°. On this thermometer there are 180° or 180 steps on the ladder of temperature.

$$1\ \text{F.}^\circ = 1.8^\circ\ \text{C.} \quad 1^\circ\ \text{C.} = \frac{1}{1.8}\ \text{F.}^\circ$$

Expressed in formulæ the changes are as follows:

$$\begin{aligned} \text{C.}^\circ &= (\text{F.}^\circ - 32^\circ)\ \tfrac{5}{9} \\ \text{F.}^\circ &= \tfrac{9}{5} \times \text{C.}^\circ + 32^\circ. \\ \text{Abs.}^\circ &= \text{C.}^\circ + 273^\circ. \end{aligned}$$

Model Examples

a. Change 42° F. to C.

$$\begin{aligned} \text{C.}^\circ &= (\text{F.}^\circ - 32)\ \tfrac{5}{9}. \\ \text{C.}^\circ &= (42 - 32)\ \tfrac{5}{9}. \\ \text{C.}^\circ &= 8 \times \tfrac{5}{9} = \tfrac{40}{9},\ 4\tfrac{4}{9}^\circ. \end{aligned}$$

$$4\tfrac{4}{9}^\circ\ \text{C.} = 42^\circ\ \text{F.}$$

b. Change 10° C. to F°.

$$\text{F.}^\circ = \tfrac{9}{5}\text{ C.} + 32^\circ.$$
$$\text{F.}^\circ = \tfrac{9}{5} \times 10 = \tfrac{90}{5}. \quad 18^\circ + 32 = 50^\circ \text{ F.}$$
$$50^\circ \text{ F.} = 10^\circ \text{ C.}$$

c. Change 26° C. to Abs.

$$\text{Abs.} = \text{C.}^\circ + 273.$$
$$\text{Abs.} = 26 + 273.$$
$$\text{Abs.} = 299^\circ.$$
$$299^\circ \text{ Abs.} = 26^\circ \text{ C.}$$

EXAMPLES

1. Change 47° F., 84° F., 97° F., −266° F., −20° F., −64° F., to C. degrees.

2. Change 87° C., 126° C., 14° C., −260° C., −123° C., 297° C., to F. degrees.

3. Change 54° C., 97° C., −273° C., −89° C., 426° C., 284° C., 54° F., 32° F. to absolute degrees.

CHARLES'S LAW

Charles's law in its numerical solution is dependent on the fact that at normal pressure gases expand $\frac{1}{273}$ of themselves for each degree Centigrade. If a gas has 100 cc. of volume at 0° degrees Centigrade (273° absolute) it will have 200 cc. at 546° absolute, which is 273° Centigrade. Under such circumstances it will have increased $\frac{273}{273}$ of itself for the whole rise in temperature, since it has risen $\frac{1}{273}$ for each degree. This regular change makes it possible for us to form a proportion showing the expansion of any gas under any change in temperature. Stated in words it follows:

The first absolute temperature: the second absolute temperature : : first volume: second volume.

We may write this in letters as follows: $T : T' : : V : V'$.

Since temperatures are regularly given in Centigrade degrees we can write this formula by taking T as a Centigrade degree and adding 273 to that. The formula would then read, $273 + T : 273 + T' : : V : V'$.

Model example: We have 1 liter of a gas at 27° C. How much shall we have at 0° C.?

Substituting in this formula given above, 273 + 27 (the first temperature): 273 (the second temperature) : : 1000 cc. (the volume of the first temperature): x (the volume at second temperature).

$$300\ x = 273000.$$
$$x = 910.$$

There will be 910 cu. cm. of gas at 0° C. if there are 1000 cc. at 27° C.

Problems

1. We have 87 cc. of oxygen at 87° C. What would we have at 64°?
2. We have 59 cc. of hydrogen at 40° C. What would we have at 92°?
3. We have 400 cc. of gaseous water at 110° C. What should we have at 128°?
4. Reduce 122 cc. of chlorine at 38° C. to N. T. (Remember that N. T. is 0° C.)
5. Reduce 244 cc. hydrochloric acid at 26° C. to N. T.
6. Reduce $2\frac{1}{2}$ liters of carbon dioxide at 18.5 to N. T.
7. We have 5.2 liters of nitrogen at 16° C. How much should we have at 42°?
8. We have 90 cc. of sulphur dioxide at 27° C. What would this become at $-273°$ C.?
9. We have 80 cc. of methane at 546°. How much should we have at 109°?
10. We have $\frac{1}{10}$ liter of methane at 37° C. How much should we have at -100?

Boyle's Law

Charles's law gives us a calculation by which we can standardize gaseous volumes and reduce them to like conditions of temperature. Boyle's law serves a similar purpose with regard to pressure. By Charles's law we reduce to N. T. 0° C. By Boyle's law we reduce to N. P. 760 millimeters or 76 centimeters of pressure.

Boyle's law states that the volume of any gas varies inversely as the pressure applied. Expressing this as a formula:

1st pressure : 2d pressure : : 2d volume : 1st volume.

or

Greater pressure : lesser pressure : : greater volume : lesser volume. $P : P' : : V' : V$.

This may be illustrated graphically by the following diagram in which the first pressure is 100 grams, the second is 200 grams, the first volume is 50 cc., the second 25 cc.

Pressure	Volume
100 gms.	50 cc.
200 gms.	25 cc.

Four quantities are inversely proportional when the quantities instead of following such an example as that shown in Charles's law, where the first temperature: second temperature: : first volume: second volume, reverse that usual arrangement and say:

The first pressure: second pressure: : second volume: first volume.

The necessity for an inverse proportion in Boyle's law is evident when one realizes that when the greater pressure is the second pressure, the second volume, which is the lesser volume, must become the first.

We have in Boyle's law an example of an inverse proportion.

FIRST MODEL PROBLEM ON BOYLE'S LAW

We have 1000 cc. of carbon dioxide gas under a pressure of 1500 grams. How much gas will there be if it has 2000 grams applied to it? In this case the first pressure is 1500 grams, the second is 2000 grams; the volume at the first pressure is 1000 cc., the volume at the second pressure is x. The formula reads:

$$P : P' : : V' : V. \quad \text{Then } 1500 : 2000 : : x : 1000.$$
$$2000\,x = 150{,}000.$$
$$x = 750.$$

SECOND MODEL PROBLEM ON BOYLE'S LAW

The volume of a gas at 78.4 cm is 1000. What will it be at normal pressure?

$$P : P' : : V' : V.$$
$$76 : 78.4 : : 1000 : x.$$
$$76x = 78400.$$
$$x = 1031.58 \text{ cc. at } 76 \text{ cm. pressure.}$$

PROBLEMS

1. We have 2800 cc. of hydrogen at a pressure of 820 millimeters, what would it be at 784 mm.?

2. We have 96 cc. of oxygen at 920 millimeters. What would we have at 440?

3. We have 86 cc. of air at 500 millimeters pressure. What shall we have at 1000?

4. We have 86 liters of sulphur dioxide at 740. What shall we have at 370?

(Remember that one atmosphere is a pressure equal to 760 millimeters of mercury.)

5. We have 840 cc. of nitrogen at $1\frac{1}{2}$ atmosphere. What shall we have at 4 atmospheres?

6. We have 96 cc. chlorine at $\frac{3}{4}$ of an atmosphere. What shall we have at $2\frac{1}{4}$ atmosphere?

7. We have 108 cc. of gaseous water at one atmosphere, what shall we have at 782 mm.?

8. Reduce to N. P. the following amounts of gases measured at 840 millimeters:

484 cc., 1286 cc., 1442 cc., 24.56 liters, 2.98 liters.

In common practice we are practically never called upon to calculate the volume of a gas which is being measured at zero degrees Centigrade. In consequence it is necessary for us to reduce the volume of a gas to both normal temperature and pressure, that is to reduce it to zero degree Centigrade and 760 millimeters of pressure from some different temperature and pressure. There is a somewhat complex formula by which if we know the volume of a gas we can determine both temperature and pressure by a single step. It is generally easier for the student to reduce any gas volume first for temperature and then for pressure, taking the answer of the first and using it in the second.

MODEL PROBLEM ON BOYLE'S AND CHARLES'S LAW

We have 100 cc. of oxygen at 780 mm. and 27° C. How much would there be at N. T. P.?

$$T : T' : : V : V'.$$

$273° : (273 + 27) = 300° : : x : 100.$

$x = 91$ cc. at 0° C. and 780 mm. pressure.

$$P : P' : : V' : V.$$

$760 : 780 : : 91 : x.$

$760\, x = 70980.$

$x = 93{,}39$ cc. at N.T.P. or 0° and 760 mm.

PROBLEMS

1. Reduce 846 cc. of oxygen at 19.5° and 764 mm. to N. T. P.
2. Reduce 2.4 liters of hydrogen at 2 atmospheres and 100° C. to N. T. P.
3. Reduce 440 cc. of nitrogen at 286° C. and 720 mm. pressure to N. T. P.
4. Reduce 4.3 liters nitrogen at ½ atmos. and −80° C. to N. T. P.
5. Reduce 4.5 liters at 22° and 757 mm. to N. T. P.

In all these calculations fixed habits of calculation will be of great service. It is difficult for the beginner to place pressures, volumes, and temperatures where they should go. It is wise in solving these problems to form the invariable habit of setting down the formula in every problem performed, and always to write down $P : P' : : V' : V$ in every Boyle's law problem, for example. It is well to remember in arranging ratios that like quantities must be grouped in the same ratio, volume with volume and temperature with temperature.

DETERMINATION OF SIMPLEST FORMULA

The determination of the simplest formula which applies to any compound is not uncommonly required in chemical arithmetic. A model example may be stated as follows.

A certain substance shows the following composition upon analysis:

		Atomic Weight.
Hydrogen	1.59 per cent	1
Nitrogen	22.22 per cent	14
Oxygen	76.17 per cent	16

If we divide these per cents by the lowest per cent we shall reduce one amount to unity.

$$1.59 \div 1.59 = 1; \quad 22.22 \div 1.59 = 13.96; \quad 76 \div 1.59 = 47.91.$$

These numbers are so close to the atomic weights given above that we may call them 1 for hydrogen, 14 for nitrogen, and 48 for oxygen. Then dividing these figures by the atomic weights of hydrogen, nitrogen, and oxygen respectively we have:

$1 \div 1 = 1$, number atoms of H.
$14 \div 14 = 1$, number atoms of N.
$48 \div 16 = 3$, number atoms of O.

Symbol is HNO_3.

A second model example follows:

A certain substance on analysis presents the following composition:

Carbon	10.04 per cent.
Hydrogen	0.84 per cent.
Chlorine	89.12 per cent.

$Cl = 35.5$. $C = 12$. $H = 1$.

Give simplest formula.

Carbon 10.04 per cent $\div 0.84 = 11.95$, practically 12.
Hydrogen $0.84 \div 0.84 = 1$.
Chlorine $89.12 \div 0.84 = 106.09$, practically 106.

Then proceeding as in the preceding problem,

$12 \div 12 = 1$ atom of C.
$1 \div 1 = 1$ atom of H.
$106 \div 35.5 = 3$ atoms of Cl.

Symbol is $CHCl_3$.

The following analyses show the composition of four common substances. Give the simplest formula of each, taking the atomic weights from the list on the cover page. Follow the two models given just above.

1. $C = 27.27$
   $O = 72.72$

2. $H = 11.11$
   $O = 88.89$

3. $N = 82.353$
   $H = 17.647$

4. $Fe = 70$
   $O = 30$

PART II

EPITOMES

CHAPTER I. According to the atomic hypothesis, a statement of a belief based on experimentation, the matter of the earth is made up of tiny systems composed of infinitely small particles. These grouped particles of matter may be considered the physical units of masses. All matter is made up of molecules which build masses. Molecules are built up from atoms. The atom acts as if it were indivisible in all ordinary chemical action.

There are between seventy and eighty different elementary atoms which cannot be separated into any simpler substance. Molecules may be either elementary or compound. The iron molecule is an elementary molecule. No element save iron can be obtained from it. The iron-rust molecule is a compound molecule, for it can be separated into three elements, iron, oxygen, and hydrogen. Less than thirty elements are available for every-day use and most of the matter in the world is made up from these thirty simple substances combined in various compounds. Elements may be divided into metals and non-metals, metals being ordinarily good conductors of heat and electricity and having a metallic luster, non-metals being ordinarily poor conductors of heat and electricity and dull in appearance. There are three states of matter, solid, liquid, and gas. Example: ice, water, and steam.

Physical change is a change in which no essential difference is made in the composition of the body suffering the change. The original substance can be recovered by physical means. In a chemical change the body itself has changed its structure and a new substance has been formed.

CHAPTER II. Our five senses, sight, smell, taste, hearing, touch, give us all our knowledge of the outside world. All our tests in chemistry are made by the use of these five senses. In the laboratory a measure of weight generally replaces touch. The balance gives us many of our most valuable chemical tests. A gain in weight often means that something has been added to a body. A loss of weight often means that a body has separated into two or more parts.

Most of the chemical changes that we meet act in one of three ways:

1. Substances may be separated into two or more parts. 2. Substances may be added together. 3. Substances may exchange the parts which compose them.

Separation of compounds is called analysis. Addition of substances is called synthesis. Exchange of parts of substances is called metathesis. Weight is our simplest means of comparison between different bodies in every-day life and in the laboratory. Standards are necessary for all comparison. There are two systems of standards; the English system, based on the pound and the yard; the metric system, based on the gram and the meter. The English system is difficult for scientific work because it is not a decimal system like our system of coinage.

Since the gram is the weight of a cubic centimeter of water at 4° C, the weight of any body of water by the metric system is practically equal to its volume. This statement is true of no other liquid save water.

Two thermometer scales are in common use, the Fahrenheit and the Centigrade. The freezing-point of water in F. is 32°, in C. is 0°. Boiling-point of water in F., 212°; in C., 100°. Density is the ratio of weight to volume. The varying density of different bodies gives us a clear method of comparison between them. Density is defined as the amount of matter per unit of volume. The collection of gases is largely dependent upon their density. To bring about chemical changes we may employ heat, light, electricity, the simple union of substances, one or all of them.

CHAPTER III. OXYGEN. — Oxygen is the gas which keeps the fires of the world burning. The atmosphere is made up of approximately one fifth oxygen and four fifths nitrogen by volume. Burning is commonly the union of some substance with oxygen. When metal burns the oxides are heavier than the metals. When wood or coal burns the oxides are heavier at the end than at the beginning. A large number of the oxides so formed are heavy invisible gases which go off through the chimney. All substances burned in oxygen become heavier because they add to themselves oxygen from the air.

Priestley, the discoverer of oxygen, first prepared the gas by burning mercury to form mercury oxide, and then heating this oxide strongly until it broke down, giving mercury and oxygen. In the laboratory we prepare the gas from a mixture of potassium chlorate and manganese dioxide. The potassium chlorate furnishes the gas, the manganese dioxide acts as a catalytic agent. A catalytic agent is a substance which, by its presence, enables an action to be more readily carried on, but which comes out of the action in the same form in which it enters. The test for oxygen is the relighting of a glowing splinter. Oxygen is insoluble, colorless, tasteless, and odorless, slightly heavier than air. Substances burning in oxygen or in air form compounds of oxygen, oxides.

Most of the earth's crust is made up of oxides. Iron-rust and iron oxide are formed by oxidation. Carbon dioxide gas is formed by the oxidation of carbon in ordinary fuel. Many processes of daily life are examples of oxidation. Breathing, respiration, decay, and changes of fermentation such as the change of cider to vinegar, are examples of oxidation. Buildings are made fireproof by using in their construction substances which will not burn. Such substances are largely compounds of oxygen, because such compounds, in general, will not unite with the free oxygen of the air.

To kindle flame a series of ignitions is commonly necessary. In a match phosphorus kindles sulphur and sulphur kindles wood. In starting a fire a match kindles paper and paper kindles wood.

CHAPTER IV. HYDROGEN. — When burned in oxygen hydrogen forms water (hydrogen oxide). Hydrogen is rarely found free in nature, but its oxide, water, covers more than half the world. Lightest of all gases it serves as a universal standard of atomic weight. The weight of a liter of hydrogen is .09 grams at standard temperature and pressure. Air is 14.4 times as heavy as hydrogen. Hydrogen is present in all acids, giving them their acid properties. It is obtained from the action of a metal on an acid, the metal replacing the hydrogen of the acid and the gas going off. Hydrogen is commonly prepared by adding zinc to either sulphuric or hydrochloric acid. In this action hydrogen goes off and leaves zinc sulphate or zinc chloride behind.

Hydrogen is lighter than air, insoluble, colorless, odorless, and tasteless. It does not support combustion, but burns freely in oxygen with a blue flame, forming hydrogen oxide (water). A mixture of hydrogen and oxygen is explosive when ignited, so great care must be taken in lighting hydrogen. Beside water, which contains one atom of oxygen, hydrogen has another oxide which contains two atoms of oxygen, hydrogen dioxide, or peroxide, a liquid which readily gives up one of its oxygen atoms, leaving water behind. Hydrogen dioxide is a powerful disinfectant, bleach, and renovator. It is made in the laboratory by treating the dioxide of the metal barium with sulphuric or hydrochloric acid.

The flame of hydrogen is intensely hot, so hot that it melts many metals. This oxidation of hydrogen provides heat which gives the flame of the calcium light. Hydrogen has been used as a gas for airships with considerable success. It may be prepared by passing steam (hydrogen oxide) over red-hot iron. In such a case hydrogen goes off and iron oxide is left behind.

CHAPTER V. ACID BASE AND SALT. — Vinegar, ammonia, and table salt are three examples of the acid, the base, and the salt. Three great inorganic acids are sulphuric, nitric, and hydrochloric acids, whose common names are oil of vitriol, nitric and muriatic acids. Among the other inorganic acids are found sulphurous, hyposulphurous, and nitrous acids.

All acids contain hydrogen. All of them lose their acid properties when hydrogen is removed. They are commonly written as compounds of hydrogen. Hydrogen sulphate is the name of sulphuric acid written as a hydrogen compound, and hydrogen chloride is the name for hydrochloric acid written in the same way. All acids contain some non-metal beside hydrogen. Oxygen is commonly present in the acids, but not always. Acids generally decompose carbonates and are neutralized by them. Common examples of this action are found in gingerbread making and in the mixing of seidlitz powders.

Acids stain and rot out fabrics. They affect vegetable colors, turn blue litmus red, methyl orange red, and affect other indicators. They neutralize the class most opposed to them, the bases. Bases neutralize acids. The sting of the bee which holds formic acid is neutralized by ammonia. Acid milk is neutralized by lime water. Bases turn red litmus blue, and red methyl orange, yellow. This is an exactly opposite action to that of the acids. Bases turn turmeric red and turn phenolphthalein pink. Acids turn pink phenolphthalein white. True bases regularly contain a metal and the radical hydroxyl. Hydroxyl is a radical (a group of atoms acting as a single atom) made up of one atom oxygen and one atom hydrogen.

Bases feel slimy to the touch and almost invariably contain a metal. Acids are regularly composed of non-metals. When acids are added to bases each neutralizes the other, and a neutral substance, a normal salt, is formed. A typical example may be found in the action of hydrochloric acid on sodium hydroxide. When these two react, sodium chloride (common salt), a neutral, is formed, and water. An acid plus a base gives a salt plus water. Acid salts are those in which all the hydrogen has not been replaced by a metal. Basic salts contain the radical hydroxyl in addition to a metal and the characteristic radical or element of the salt. Many salts have a salty taste similar to that of common salt. Salts are used widely in commercial processes.

CHAPTER VI. WATER. — Water is the most wide-spread compound on the face of the earth. By its power of taking solid matter into solution and suspension it is constantly changing the surface of the earth. It is a necessity to plant life, as it furnishes plants with foods through their hair-like root tubes which draw up the nourishment of the soil in water solution. Water makes up a large percentage of the total weight of the foods that we eat. It cleanses the human frame inside and out. Pure water does not exist in nature because of the impurities which it dissolves as it comes to the soil and passes over the soil. From such solution of impurities comes the danger of water-borne diseases, such as typhoid and cholera. No health necessity is more important than to keep the water supply free from contamination. Sandy soils filter water far better than clay soils; most city water is filtered artificially in one way or another.

When crystals contain water that evaporates at the ordinary temperature of the air, this action, which leaves the solid turned to powder, is called efflorescence. Dry substances which take up water from the atmosphere and thereby become moist are called deliquescent.

Mineral water in general is formed by water trickling through beds of mineral salts. Hard water contains mineral salts in solutions. When these salts are carbonates the water is temporarily hard. It can be softened by boiling. When sulphates are present it is called permanently hard. It cannot in general be softened by boiling. Soap will not go into solution in hard water. Pure water may be obtained by distillation, by changing liquid water into gaseous water and then condensing the gas back to a liquid again.

We may prove that water is made up of hydrogen and oxygen in the proportion of two parts hydrogen to one of oxygen by the analysis of water by means of the electric current, or by synthesis, by the burning of hydrogen in oxygen with the heat of the electric spark. By both processes two volumes of hydrogen are proved to unite with one of oxygen.

The action of a steam engine depends first on the power of carbon to unite with oxygen, giving off heat. Second, on the heat generated, turning water into steam, and third, on the expansive power of steam.

Most substances expand when heated, contract when chilled. Water, cooling, contracts to a point about four degrees above the temperature of freezing water. Cooling below that point it begins to expand regularly to zero. When it turns to ice, water expands still more. Life in the temperate zone depends largely on this expansion for its continued existence. Because of this property of water the lakes and rivers never totally freeze.

CHAPTER VII. Laws. — The laws of man are dependent upon the characteristics of man. The laws of nature are infallible and unerring. They must be obeyed. Some of the great laws which govern chemical action are stated here. The law of the conservation of energy states that energy can never be destroyed; that in any separate action, just so much energy as is put in, just so much energy is taken out. The law of the conservation of matter states that matter can never be destroyed: that just so much matter as enters into any separate action, just so much matter goes out. The law of definite proportions by weight states that the same compounds always contain the same elements combined in the same proportions by weight. The law of multiple proportions by weight states that when the same two elements combine to form more than one compound, the different weights of one of the elements which unite with a constant weight of the other bear a simple ratio to one another. The law of reciprocal proportions states that the weights of different elements which combine separately with one and the same weight of another element are either the same as, or are simple multiples of, the weights of these different multiples which combine with each other.

The atomic theory is closely related to these last three great laws, and chemical combination is supposed to take place between atoms which act as if they are indivisible. Atoms have a relatively fixed weight calculated from the weight of the hydrogen atom taken as the standard. The weight of a hydrogen atom is called one microcrith. Oxygen, which is sixteen times as heavy as hydrogen, is said to weigh 16 microcriths.

CHAPTER VIII. CHLORINE. — Chlorine is the bleacher. The bleaching power of this element is due to the power which chlorine possesses of uniting with hydrogen. If you wet a cloth in water and expose it to chlorine, the chlorine unites with the hydrogen of water, setting free oxygen which burns out the color. Chlorine therefore bleaches indirectly. The bleaching of goods in the laundry is commonly due to the use of chlorine washing powders. Never found free in nature, chlorine is found everywhere in its compounds, the chlorides. Some of these chlorides are common salt, bleaching powder, and potassium chloride. Hydrochloric acid, one of the strong acids, is made up of hydrogen and chlorine.

When the electric current passes through a solution of hydrochloric acid, the acid breaks down into hydrogen and chlorine. The gas has a strangling, disagreeable smell. It is two and one-half times heavier than air, is soluble and has a yellow-green color. In the laboratory, chlorine is generally prepared by acting on hydrochloric acid with manganese dioxide. In this action the oxygen of the oxide unites with the hydrogen of the hydrochloric acid to form water, and chlorine goes free. Chlorine may be prepared also by the use of salt, sulphuric acid, and manganese dioxide.

Dry litmus paper or dry textiles are unaffected by chlorine. Wet litmus and wet colored cloths are both bleached by the gas. A solution of chlorine in water changes rapidly in the light because the chlorine joins with the hydrogen of the water to form hydrochloric acid. All plant life dies in chlorine gas because the chlorine unites with the hydrogen of the water of the plant. Its action on the human body is dangerous. In the manufacture of chlorine the gas is absorbed in large quantities in lime, forming chloride of lime, or bleaching powder. Bleaching powder gives up its chlorine when a weak acid is added to it. It then serves as a bleach.

Hydrochloric acid, commonly known as muriatic acid, is one of the great acids of the commercial world. It is found free to some degree in volcanoes and is commonly prepared from salt and sulphuric acid.

Sodium chloride + hydrogen sulphate = hydrogen chloride + sodium sulphate.

Hydrochloric acid gas is colorless. It is yellow in solution because of impurities. It is heavier than air, and a powerful solvent. It is dangerous to inhale. The gas is extremely soluble in water, five hundred volumes of gas dissolving in one volume of water. The manufacturing process by which it is prepared is similar to the laboratory process, but large quantities are produced as a by-product by a process known as the Le Blanc process. Chlorine may be tested by the addition of silver nitrate.

Silver nitrate + hydrogen chloride = silver chloride + hydrogen nitrate.

CHAPTER IX. THE CHLORIDES. — The sea salts are chiefly chlorides. Of these, sodium chloride, common salt, makes up the greater part. A gallon of sea water contains about 2500 grains of various salts, and among the metals from which these compounds are formed are potassium, magnesium, sodium, traces of gold and other rare elements. American salt chiefly comes from salt springs. European salt is obtained from rock salt mines. Absolutely pure salt is seldom found on the table. Chlorides are salts of hydrochloric acid and form a group among themselves, as do the sulphates, nitrates, etc. Calcium chloride is a deliquescent salt whose use depends largely on its capacity for seizing water from the air or other gases. Cobalt chloride is practically colorless when it contains water of crystallization, and is blue when dry. A sympathetic ink can be made by taking advantage of this characteristic.

Calomel (mercurous chloride) is a valuable medicine. Corrosive sublimate (mercuric chloride) is a valuable antiseptic and is a deadly poison when taken internally. Potassium chloride and magnesium chloride are used to prepare metallic potassium and magnesium. Gold and silver chloride are photographic salts which break down under the influence of the rays of the sun.

Tin chlorides are used for dyeing. Zinc chlorides are used for solder. Ammonium chloride (sal ammoniac) is used in electric batteries. Barium chloride is used in analytical work. Copper chloride is used in making paints. Iron perchloride is used to supply iron in the blood.

The chief test for the chlorides is the precipitate of silver chloride which solutions of these salts form on the addition of silver nitrate. One of the chief groups of metals, known in analytical chemistry as the silver group, depends on a reaction of this type for its separation.

CHAPTER X. FLUORINE, BROMINE, AND IODINE. — The halogens, fluorine, bromine, chlorine, and iodine, form a family which are very closely related to each other. Every element is more or less related to a definite group of other elements, but none are so closely interrelated as the halogens.

Fluorine is the most active of the halogens. It is used as an etcher of glass. It combines with almost all of the metals and even water breaks down in its presence. Its commonest compounds are fluorspar (calcium fluoride) and hydrofluoric acid (hydrogen fluoride). The latter is prepared by acting upon the former. Fluorine can be obtained by breaking down hydrofluoric acid with the electric current. The gas is of a greenish yellow color and is heavier than air, though its density is the least of the four halogen densities. It is held in containers of platinum and iridium.

Bromine is found in saline springs. It is a liquid at high temperatures and has a very irritating odor. It is soluble in water, boils easily, and burns the flesh. It forms hydrobromic acid whose salts are the bromides. Bromine is heavier than chlorine, as chlorine is heavier than fluorine. Iodine is the heaviest of all the group and is the only solid. Iodine is found chiefly in the kelp of the shore. It is a blue-black, crystalline solid. Iodine melts easily and its vapor is a deep indigo blue. It is an element which sublimates, turning from a gas to a solid with no apparent intervening liquid change. It is slightly soluble in water, very soluble in alcohol. Iodine is tested by the blue color which it imparts to starch. Iodine is used in medicine as a counter-irritant.

Each member of the halogen family forms an acid with hydrogen. Each forms similar salts. The energy of each of the four varies inversely as their atomic weight.. The heavier the element the less its energy.

CHAPTER XI. Carbon Dioxide. — Carbon dioxide, a product of fire and decay, acts as a food for nourishing plant life. It is found throughout the atmosphere in varying proportions.

Carbon dioxide is prepared by roasting calcium carbonate or by acting upon calcium carbonate with hydrochloric acid. When sulphuric acid is used instead of hydrochloric acid but little gas is obtained, unless the calcium carbonate is powdered, as insoluble calcium sulphate is formed.

Carbon dioxide is very heavy, easily soluble, invisible, odorless, and tasteless. The carbon dioxide escaping from the city is carried in part by the winds, but a natural means provided for its absorption is found in the trees and shrubs which take it in, break it down into carbon and oxygen, and use the carbon for food. The "cycle of carbon" is that change by which carbon turning to carbon dioxide is absorbed by the plants which separate the carbon from the oxygen and set free the oxygen. The carbon of the plant then returns to carbon dioxide either by serving as fuel or by decay.

Carbon dioxide is used commercially in mineral waters, in the soda fountain, and in fire extinguishers.

The test for the gas is the precipitate of calcium carbonate formed when carbon dioxide is run into calcium hydroxide (lime water).

Carbon dioxide raises bread and soda biscuits and is given off in fermentation.

Carbon monoxide is found in coal fires. It readily burns with a blue flame, forming carbon dioxide. This gas is used for reducing iron ore to metallic iron. It is also used in making water gas, which is prepared by passing steam over red-hot carbon. Carbon monoxide is prepared in the laboratory by acting on oxalic acid with sulphuric acid. It is tested by its forming carbon dioxide when burned.

CHAPTER XII. CARBONATES. — The carbonates are the salts of carbonic acid. Common carbonates are those of calcium, sodium, and potassium. Calcium carbonate is found in marble, coral, and limestone. This last calcium carbonate provides the lime which is so essential to building. Washing soda is the normal carbonate of sodium. Baking soda is the acid carbonate of sodium. These carbonates are found largely in lakes containing chemicals, such as the Great Salt Lake of Utah. Potassium carbonate is found in the soil broken down from granite rocks, and in the trees which take it from the soil. The carbonates all give off carbon dioxide gas when attacked by an acid. This is the chief test for the group.

Sodium carbonate is used largely in commerce and is prepared by one of two processes, the Le Blanc or the Solvay. The Le Blanc is a very old process and depends upon the following principles: Salt and sulphuric acid are first mixed together; hydrochloric acid is given off by this action and sodium sulphates are left behind. The sodium sulphates formed are then heated with lime and coal. The sodium unites with carbon dioxide given off by the coal and limestone. Impure sodium carbonate is left after the action. This, heated with coal, gives a comparatively pure carbonate.

Carbonates are commonly similar in appearance when crushed, appearing as white powders. They have a slightly alkaline action. Mortar is made by mixing unslaked lime (calcium oxide) with water. This action produces slaked lime (calcium hydroxide), which, mixed with sand, forms mortar. Sodium carbonate is largely used in glass-making. Chalk, the chalk of the schoolroom, is calcium carbonate.

CHAPTER XIII. EQUATIONS. — Chemical affinity may be defined as the attraction which one chemical body possesses for another. Every element has a symbol which represents a single atom of that element. This symbol represents also the atom's weight. The weight of an atom of hydrogen is taken as the standard of atomic weight and is called a microcrith. O, the symbol of oxygen, represents an atom sixteen times as heavy as an atom of hydrogen. O_2 represents a molecule of oxygen which contains two atoms. SO_2, the formula for sulphur dioxide gas, represents a molecule containing one atom of sulphur weighing thirty-two microcriths and two atoms of oxygen each of which weighs sixteen microcriths and which both together weigh thirty-two microcriths. The total weight of a molecule of sulphur dioxide, therefore, is sixty-four microcriths.

All reactions represent parts by weight of the molecules represented, as well as the substances themselves. Many of the symbols come from the first letter of the names of the elements, as N for nitrogen, O for oxygen. Some come from the first letters of the Latin names, as Au for gold, and Na for sodium.

The reaction $CaCO_3$ (heated) $= CaO + CO_2$ may be written either with the sign of equality as used in this book, or with arrows as written here. $CaCO_3$ (heated) $\rightleftarrows CaO + CO_2$. This second form indicates an incomplete reaction which takes place to some degree in both directions.

Three classes of equations are considered in this book, the addition or synthetical, the division or analytical, and the exchange or metathetical. A type of the first is shown in ($S + O_2 = SO_2$), of the second in (HgO (heated) $= Hg + O$) and of the third in ($AgNO_3 + NaCl = AgCl + NaNO_3$). Chemical affinity may be roughly measured by the amount of energy set free. Most of the reactions which we have studied take place between substances in solution which break down into positively and negatively charged parts called ions. Ions are electrically charged atoms or radicals. Bodies charged with like electricity repel, bodies charged with unlike electricity attract. Positive ions attract negative ions. From the interchange of positive and negative ions in conducting solutions come chemical changes in general. Substances which come out of solution in reactions, such as precipitates or gases, are supposed to generally lose their external charge when they leave the solution. The more insoluble the compound, the more likely it is to form and come out of the solution.

Valence may be defined as the combining power which a radical or an elementary atom possesses in regard to hydrogen or an equivalent of hydrogen. A monovalent substance combines with or replaces one atom of hydrogen, a bivalent substance combines with or replaces two atoms of hydrogen, a trivalent substance combines with or replaces three.

CHAPTER XIV. NITROGEN.—Nitrogen is the restraining gas of the air. It can be made to unite with other elements only with great difficulty. It is found largely in foods and is an essential part of the body-building proteids. It appears in the decay of many organic compounds in the form of ammonia, of nitrates and nitrites. Nitrogen may be prepared in the laboratory by burning out the oxygen of the air with phosphorus and collecting the nitrogen left behind, by heating ammonia nitrite or by heating ammonium chloride and sodium nitrite together. This reaction produces ammonium nitrite, which breaks up yielding nitrogen. Nitrogen is very slightly lighter than air, puts out a flame and is commonly recognized by its inertness.

Ammonia water is the old spirits of hartshorn. Ammonia gas is a common product of decay. Its presence to any extent in spring water is often a sign of danger. Nitrogen in foods comes from the nitrogen in combination in the atmosphere and in the soil.

Many fertilizers depend upon their content of ammonia for their effectiveness, as the uncombined nitrogen of the air cannot be used directly by most plants. Ammonia is found in nature in volcanic lands. It can be prepared in the laboratory by combining nitrogen and hydrogen with the electric spark, or by heating ammonia water which is prepared from the ammoniacal liquid of the gas works. Ammonia can be tested by its peculiar pungent odor and its alkaline action on litmus paper. It is used commercially in cold storage and in the preparation of artificial ice. Ammonia water is used as a cleanser, in dyeing and in the arts. Ammonia gas run into water produces ammonium hydroxide. Ammonia water produced in this way is the ordinary household ammonia.

The oxides of nitrogen form a group of considerable interest. Nitrous oxide contains two atoms of nitrogen and one of oxygen. It is an anæsthetic and has a slightly sweetish taste and smell. Nitric oxide has one atom of nitrogen and one of oxygen and is made by the action of nitric acid on copper. It is a colorless, slightly soluble gas. Nitrogen peroxide is made by adding oxygen to nitric oxide. It is a red-brown gas which contains two atoms of nitrogen to four of oxygen. It is very poisonous and suffocating. Nitrogen pentoxide is chiefly of interest as the theoretical basis of nitric acid. When added to water it forms nitric acid. It is prepared by removing water from nitric acid. It contains two atoms of nitrogen and five of oxygen.

CHAPTER XV. NITRIC ACID. — Nitric acid is the chief solvent of the metals. It is found to some slight degree free in the air, but its salts, the nitrates, are found widely as the product of organic change. Chemists now prepare nitrates from the oxygen of the air by burning the nitrogen of the air in oxygen, forming nitric acid and neutralizing this acid with calcium hydroxide, forming calcium nitrate. Nitric acid is prepared commercially by heating some nitrate like sodium nitrate and sulphuric acid together, producing sodium sulphate and nitric acid (hydrogen nitrate).

This acid is very corrosive, colorless, and not especially stable in composition. It is a great oxidizing agent. It forms sulphuric acid by oxidizing sulphur to sulphur trioxide when it is boiled with sulphur. We can test nitric acid by the blue color of copper nitrate formed when the acid is heated with copper, and by the brown ring of nitrogen peroxide which forms when ferrous sulphate is poured on the acid in the tube.

When nitrates are to be tested, concentrated sulphuric acid must be added to the salt in order that free nitric acid may be given off. Nitric acid when joined with hydrochloric acid dissolves gold. The two acids form aqua regia, a powerful solvent. Nitrates are used largely in explosives. Nitrogen is used in the structure of one group of plants, the legumes, which, by the aid of certain bacteria, are enabled to build the nitrogen of the air into their structure.

CHAPTER XVI. AIR. — We live at the bottom of an ocean of air which extends upward for many miles. The density of this ocean of air is constantly varying as we rise into it, from the comparatively high density of surface air to the rarest matter we can conceive. Such varying density of the air is due to its great compressibility. The density of the air at the surface of the earth is constantly changing from a similar cause. The air contains about four fifths nitrogen and one fifth oxygen by volume. This may be shown by the burning of phosphorus in air to form phosphorous oxide. Oxygen is the great combiner of the air. Nitrogen is the gas which holds oxygen in check. The inert elementary gas argon, one of the constituents of air, was found in comparatively recent years by Ramsay and Rayleigh. Argon is even more inert than nitrogen and it does not combine with any known substance.

Water vapor is always present in air. The presence of water vapor accounts for rain, dew, snow, and hail, for the humidity of the dog-day as well as for the snow of winter. The various gases of the atmosphere are of different weights, but they keep an even mixture by means of their property of diffusion. Air can be liquefied and solidified by the use of extreme cold and pressure. Air is a mixture and not a compound. Every constituent of air acts in air as it does out of air. Oxygen forms oxides, nitrogen combines with magnesium, carbon dioxide gives the same test with lime water and produces white calcium carbonate in air or out of air.

CHAPTER XVII. THE GAS LAWS. — All gases contract regularly with cold and expand with heat. The law of Charles or of Dalton states that every true gas expands or contracts by a certain definite percentage of its volume for a rise or fall of one degree Centigrade. This expansion or contraction is practically $\frac{1}{273}$ of the volume of the gas at zero degrees Centigrade. Theoretically, if gases could reach -273 degrees Centigrade they would disappear, a condition which is impossible because of the law of the conservation of matter. Every known gas becomes a liquid before -273 degrees Centigrade is reached. This point, -273 degrees Centigrade is considered the starting-point of temperature and is called absolute zero. N.T. or S.T. is zero degrees Centigrade or -273 absolute. Gas volumes must be reduced to this temperature in order that they may be compared with other gases under like conditions.

The mercury in the barometer tube rises and falls as the weight of the air increases or decreases. The pressure of the air like its temperature is constantly varying. The lighter the atmosphere the more the gas expands. The heavier the atmosphere the more the gas contracts. Gases expand or contract readily under change of pressure. The law of Boyle states that the volume of every true gas varies inversely as the pressure that is put upon it. Stating this in proportion: The first or greater pressure : second or smaller pressure : : second or large volume : the first or smaller volume. Stating this in another fashion, the greater the pressure exerted on any gas the smaller its volume will become. Gas volumes are compared at a normal or standard pressure which is taken as 760 millimeters of mercury. The letters N.T.P. or S.T.P. stand for a pressure equal to 760 millimeters of mercury and a temperature of zero degrees Centigrade or 273 degrees absolute.

The law of Gay Lussac states that gases combine in volumes which bear a simple ratio to each other and to the volumes of the product obtained. For example, 2 volumes of hydrogen + 1 volume of oxygen = 2 volumes of gaseous water. Avogadro's hypothesis is largely based upon these gas laws. This hypothesis states that the same number of molecules may be found in equal volumes of gases at the same temperature and pressure. On this theory we base a large number of our beliefs to-day.

CHAPTER XVIII. Soil. — There were four elements in the natural philosophy of the ancients, air, earth, fire, and water. None of these substances were really elementary. The earth is made up of a few elements and of thousands and thousands of compounds built from metals and non-metals. From the compounds of the soil come food for man and beast. Plants generally obtain the elements of their structure (other than hydrogen, carbon, and oxygen) from the soil. As the plants use elements, the soil to be fertile must contain the elements required by the plants. These elements are supplied by manures. Nitrates give nitrogen; guano, phosphates; barnyard manure gives both; lime gives calcium and neutralizes the acids of the soil. The rotation of crops sometimes supplies necessary plant food.

The soil supports plant life mechanically as well as chemically. Earth is broken-down rock disintegrated largely by the weathering. Silicon dioxide (sand), one of the chief constituents of the earth, is found almost everywhere, not only as sand but as quartz, flint, amethyst, and jasper. Silica is practically insoluble in water, is slightly soluble in alkalies, and is found in the coating of various grasses and other plants. Clay contains silica and alumina or oxide of aluminum. Aluminum is seldom found in the plant life which rises above the earth. Granite is made up of quartz which is silica, feldspar which contains silica, aluminum, and potassium, and mica which contains these three in slightly different form. Oxides of iron and magnesium are also found in granite. Granite broken down leaves clay behind.

CHAPTER XIX. Sulphur. — Sulphur is found largely in southern

Italy, especially in Sicily, in Japan and the New Hebrides, in the western part of the United States, and in Mexico. Sulphur appears commonly in strata which occur between other strata of clay. It is especially likely to occur, however, in volcanic lands. Sulphur combines readily with silver or copper and tarnishes both metals. Paints made of white lead are easily tarnished or blackened by sulphur. Sulphur is found in eggs, vegetables, and in the human body. This element acts like oxygen in many ways and is a great combiner. Among ores in which sulphur is found are cinnabar (mercury sulphide) and iron pyrites (iron sulphide). The sulphur obtained from iron pyrites is chiefly of importance in the production of sulphuric acid. Sulphur is a light solid at ordinary temperatures, melts readily and vaporizes easily, is insoluble in water and soluble in carbon bisulphide.

Sulphur occurs in four allotropic forms, prismatic, rhombic, plastic, and white amorphous sulphur. When found in nature, sulphur is almost always mixed with earthy matters. Sulphur is prepared by heating the earthy mass, with which it is mixed, out of contact with the air. Sulphur is used in rubber manufacture, in gunpowder, in matches, in acids, and as a disinfectant. In this last capacity it is used to clean vineyards from attacking fungi.

CHAPTER XX. SULPHUR DIOXIDE, SULPHUR TRIOXIDE, AND HYDROGEN SULPHIDE. — Sulphur dioxide is largely used as a mild bleach. This gas is produced by burning sulphur in oxygen or in the air. It allows no oxidation in its presence and in consequence destroys germ life which requires oxygen. It is used in the making of artificial ice and in preparing pulp for paper. Sulphur dioxide can be prepared from the action of heated sulphuric acid on copper. It bleaches vegetable colors like those which tinge the flowers, fading them to a dull white.

Sulphur trioxide is of chief importance as the basis of sulphuric acid. When water is added to sulphur trioxide, sulphuric acid is formed. Such formation of acids is characteristic of the non-metallic gaseous oxides. Non-metallic acid-forming oxides are called acid anhydrides. We have sulphurous anhydride, which is sulphur dioxide, and sulphuric anhydride, which is sulphur trioxide. The ending ous commonly indicates a lower form or one which contains less matter than some other similar chemical compound, made up of the same elements but containing more atoms in the molecule. The ending ic commonly denotes such a higher form. In sulphurous acid we find that the two hydrogens are separated from the sulphur and the three oxygens which compose the acid. In this acid we find the sulphur atom and three oxygen atoms, joined together, form a group of atoms acting as a single atom which may be called a radical. Sulphurous acid has a radical composed of one atom of sulphur and three atoms of oxygen corresponding to the ending ous. We find that the radical of sulphurous acid forms compounds known as sulphites. The radical of sulphuric acid forms compounds known as sulphates. Sulphur dioxide can be turned to sulphur trioxide by the catalytic action which occurs when sulphurous anhydride is passed over red-hot platinum. Under such circumstances the sulphur dioxide unites with the oxygen of the air and forms sulphur trioxide. Sulphur trioxide (sulphuric anhydride) is a hygroscopic body. Hydrogen sulphide, known commonly as sulphuretted hydrogen and sometimes called hydrosulphuric acid, is the gas which comes from rotten eggs, rises from poor coal which contains sulphur and is found in many gas flames. It is found in nature in volcanic surroundings. Hydrogen sulphide readily combines with metals, and tarnishes copper and silver. It is used largely in analytical separations of the metals. A whole group of metals, mercury, arsenic, lead, tin, bismuth, antimony, copper, and cadmium is precipitated from solutions which contain them by means of this reagent. Sulphuretted hydrogen is slightly poisonous and should be handled with considerable caution.

CHAPTER XXI. SULPHUROUS ACID AND THE SULPHITES, SULPHURIC ACID AND THE SULPHATES. — Sulphurous acid is of special value in paper-making. Its compounds, the sulphites, are used for most grades of paper in which wood pulp is used. Wood pulp is prepared from wood by treatment with sulphite baths. Paper-making depends largely upon the fibrous quality of its constituent parts. Wood pulp which has not been chemically treated is used by itself in paper-making but not for manufacturing strong paper. When wood pulp is treated with the sulphite bath the fibers of the wood are strengthened and bleached. Farther bleaching is obtained from chlorine. Good papers generally contain a fair percentage of sulphite pulp. In the manufacture of paper the wood pulp is thrown into a beating machine and broken up. As it is bleached and felted, dried and polished it turns to paper. This process is dependent to a large degree on the power which certain chemicals possess of strengthening wood fibers.

Sulphurous acid or hydrogen sulphite forms compounds, sulphites. Sulphites may be tested by the addition of barium chloride. When this substance is added a white precipitate falls which is easily soluble in hydrochloric acid. Sulphurous acid is a powerful reducing agent readily changing to sulphuric acid.

Sulphuric acid is one of the greatest commercial necessities of modern life, is used in electric generators and in storage batteries in textile work, in the production of grape sugar and fertilizers, in the purification of oils and in the forming of other acids. It may be produced by running sulphur trioxide into water. Sulphuric acid readily dissolves most of the metals with the evolution of hydrogen. It is especially dangerous when it acts in its concentrated form on human flesh. This acid, like most others, is readily broken down by a current of electricity. In mixing sulphuric acid and water the acid should always be poured on water. Water should never be poured on the acid.

There are two methods for producing sulphuric acid, the contact or catalytic process and the English or chamber process. In the contact process sulphur dioxide is obtained by burning sulphur or roasting iron sulphide. The gas obtained is then passed over hot platinum. Sulphur trioxide is formed by catalytic action, and sulphuric acid is produced by running the trioxide into water. When sulphur trioxide joins water, sulphuric acid is produced. In the English or chamber process sulphur dioxide, produced as in the contact process, is run into lead chambers where it is mixed with moist air and oxides of nitrogen. In these chambers it changes into sulphur trioxide. The trioxide formed unites immediately with steam and passes out as sulphuric acid.

Sulphuric acid, from its hygroscopic characteristics, serves many purposes. Moist gases passed through this liquid come out dried, leaving their water behind. Oils may be separated from water by taking advantage of the same property of the acid. The production of other acids by means of sulphuric acid depends largely on the fact that sulphuric acid boils (vaporizes) at a higher temperature than most of the other acids.

Sulphates are formed by the action of sulphuric acid upon the metals. The test for sulphuric acid is similar to that used for testing sulphurous acid. Barium chloride forms a white precipitate, barium sulphate, when added to sulphuric acid. Barium sulphate differs from the barium sulphite formed with sulphurous acid by its insolubility in hydrochloric acid.

CHAPTER XXII. PHOSPHORUS AND PHOSPHATES. — Phosphorus, the most necessary element of a match, is used for this purpose because it burns so readily. This element must be kept in water in order to keep it from oxidizing. Two allotropes of phosphorus exist, the yellow and the red. The first oxidizes very easily, the second oxidizes only with difficulty. The first is poisonous, the second is not. The first burns with an odor, the second is odorless. The first is used for the common match, the second for safety matches. Friction matches cause many fires. They are commonly made of phosphorus, sulphur, some oxidizing material, and wood. The use of safety matches depends largely on the property which red phosphorus possesses of changing to yellow phosphorus by aid of the heat of friction. Phosphorus is a necessary requirement for fertile fields. Phosphates are largely used in baking powders in the form of an acid salt which, added to baking soda, produces carbon dioxide gas which raises bread.

CHAPTER XXIII. THE THEORY OF SOLUTION AND THE EFFECT OF ELECTRICITY ON CHEMICAL ACTION. — The electric current is generally produced by the transformation of heat or chemical energy turned into electrical energy. Electricity may be passed from a positive terminal or electrode to a negative terminal or electrode. This is true not only of two terminals separated by air, such as we see in the arc light, but also of electricity passing between two terminals immersed in a solution. When an electric current passes through a conducting solution marked changes occur. The ions of the solution tend to separate. The positive ions are attracted to the negative pole, the negative ions to the positive pole. Solutions which conduct electricity and are broken down by the current as it passes are called electrolytes. Solutions which will not pass the current are called non-electrolytes. In any solution of an electrolyte, we believe there are present both electrically charged ions (which are produced by the dissociation of larger electrically neutral molecules) and undissociated molecules. There is a constant interchange of parts between the dissociated ions and the undissociated molecules. Valence, according to the ionic theory, becomes a matter of the number of positive and negative charges carried by an ion. Monovalent substances carry one positive or negative charge, bivalent two, and trivalent three. Small plus and minus signs indicate the ionic form. The reaction $Na^+(OH)^- + H^+Cl^- = Na^+Cl^- + H_2O$ indicates that the first three substances are in the ionic form and the fourth is in the molecular form.

The positive pole which conducts electricity in the electrolyte is called an anode. The negative is called a cathode. The negative ion which seeks the anode is called an anion, and the positive ion which seeks the negative pole is called a cation. Electroplating is dependent upon the fact that ions when they reach a terminal in the solution, lose their electrical charge and deposit on that terminal. According to Faraday's law, the ratio between the quantities of separated matter obtained by electrolysis is the same as that which exists between their chemical equivalents. The chemical equivalent of an element is its atomic weight divided by its valence.

CHAPTER XXIV. Carbon. — Carbon is found in several forms, as diamond, graphite, lampblack, charcoal, coke, gas, and coal. There are three true allotropes of carbon, diamond, graphite, and amorphous carbon. The diamond is pure carbon and forms carbon dioxide when burned. Diamonds are found mainly in South Africa and exist in meteorites. Henri Moissan was the producer of the artificial diamond. Diamonds exist in various colors, are crystalline, are very hard, and have wonderful properties of transparency and reflection. They are used for drills, glass cutters, and rock borers, as well as for jewels.

Graphite is the lead of the lead pencil, the protective coating of iron, and lubricator of machinery. This form of carbon is used in making gunpowder and crucibles and is an aid in electrical processes. Lead pencils are made of clean graphite mixed with clean clay. Graphite forms carbon dioxide when burned and is non-crystalline in structure. It is found largely in Siberia and California. There are a number of different carbon forms in the group of amorphous carbon, such as lampblack and wood charcoal. This last-named substance is produced by heating wood out of contact with the air in order that carbon dioxide may not be formed by union with oxygen. Charcoal is used largely as a filterer and deodorizer. Coal is an impure form of carbon. Coals may be divided into bituminous (soft) and anthracite (hard). Lignite is a coal which contains less carbon than bituminous. Peat comes from the bogs and contains less carbon than lignite. Coal is produced by the gradual decomposition of decaying organic matter under the influence of natural forces.

The cyanides are poisonous inorganic salts. The acid from which they come, prussic acid or hydrogen cyanide, is especially poisonous. Potassium cyanide is used in photography and in reducing metals in the laboratory. The cyanides of iron are used to some degree as dyes. The radical cyanogen is made up of one atom each of nitrogen and carbon.

CHAPTER XXV. METHANE. — Organic chemistry may be defined as the chemistry of the carbon compounds. Organic and inorganic substances obey the same great chemical laws and differ chiefly in their chemical characteristics. These differences are largely due to the invariable presence of one element in organic substances, the element carbon.

Methane is a typical organic substance which is commonly known as marsh gas. It is composed of one atom of carbon and four atoms of hydrogen. The valence of carbon determines the constitution of many organic substances. Carbon is quadrivalent, that is, it commonly combines with four hydrogen atoms or their equivalent. Valence may be defined in a general way as the capacity which any element possesses of combining with or replacing hydrogen. Methane is prepared from sodium acetate and sodium hydroxide acting in the presence of quicklime. This gas burns with a hot flame, forming carbon dioxide and water. Wood alcohol, a hydroxide derivative of methane, is a valuable solvent and fuel produced by the destructive distillation of wood. The aldehyde derivatives of methane contain the aldehyde radical made up of one atom of carbon, one of hydrogen, and one of oxygen. Formaldehyde, the disinfectant, is one of the most typical aldehydes. Formic acid is the organic acid directly derived from methane. It contains the radical carboxyl made up of one atom of carbon, two of oxygen, and one of hydrogen. The oxide derivative of methane is an ether, methyl ether. Chloroform is one of the chloride derivatives of the same compound gas.

CHAPTER XXVI. ETHANE AND ITS DERIVATIVES. — The gas ethane is colorless, tasteless, and insoluble, and burns with a slightly luminous flame. Among its derivatives are anæsthetic ether, vinegar, the esters and acetates. Ethane contains two carbon atoms and six hydrogen atoms. Like methane it belongs to the series known as the hydrocarbons. Each part of this series is formed by adding a carbon and two hydrogen atoms together.

Ethyl alcohol (spirits of wine) is the hydroxide derivative of ethane. It is formed by fermentation which produces an impure alcohol. This impure alcohol is purified by fractional distillation, a process which depends upon the different boiling-points of two or more liquids.

Alcohol is a colorless liquid lighter than water, which boils at 78 degrees C. and solidifies at – 170 degrees C. It is of value in making spirit thermometers because of the low temperature at which it solidifies. It oxidizes, when burned, to carbon dioxide and water. The test for alcohol is the formation of iodoform in an alcoholic solution. Ether is the best known anæsthetic. Sulphuric ether is the oxide derivative of ethane. It is prepared by the hygroscopic action of sulphuric acid on ethyl alcohol. It is very volatile and is an inflammable substance. Great care must be used in heating it. The esters are salts of ethyl alcohol found in many fruits and flowers. Many such compounds can be prepared in the laboratory. They are commonly made by the neutralizing action of alcohol on an organic or inorganic acid.

Acetic acid is one of the liquids separated from pyroligneous acid, the acid left after the destructive distillation of wood. It is an organic acid characterized by the radical carboxyl, a radical composed of one atom of carbon and two of oxygen, and one of hydrogen. Acetic acid is formed from calcium acetate by the action of concentrated hydrochloric acid. This acid is used in making acetates and in the manufacture of white lead for paints. It forms a series of salts known as the acetates.

CHAPTER XXVII. FERMENTATION. — The garden of the air contains many tiny plants. There are three great divisions of these plants, the bacteria, yeasts, and molds. These minute organisms are present in every room in thousands, and grow swiftly where they can find three requisites, heat, moisture, and nutriment. The molds are the largest of the three divisions and grow to a considerable degree on any food, in moisture and warmth. The yeasts are best known by the yeast cake. The bacteria are the scavengers of the air, many of whom are constantly breaking down organic wastes into harmless organic matter which serves as food for plant life. While a number of the bacteria are of great assistance in the development of various necessary processes, it must be remembered that in this group are found the dangerous germs of typhoid, consumption, and pneumonia.

The three classes are commonly spoken of as micro-organisms, microbes, or germs. These micro-organisms obtain their nutriment from the food on which they lodge, rather than taking a part of it from the air as do the plants. The science of the preservation of food consists largely of shielding it from the attack of harmful organisms. From this necessity comes the salting of meats and the covering of jellies with paraffine. The changing of soft cider to hard cider is due to the growth of yeasts. Yeast growth is at the basis of fermentation. When yeasts grow in a liquid medium, carbon dioxide gas and alcohol are evolved, thence comes the alcohol of hard cider and all fermented liquors. Bread-raising is produced by the evolution of carbon dioxide gas produced when yeast grows in bread, evolving gaseous carbon dioxide. Alcohol, formed in the fermentation, is driven off by baking bread. Yeasts grow by budding. Bacteria grow by lengthening and division. Acetic acid is the acid principle of vinegar. It is produced by a bacterial action on the alcohol of hard cider which results in oxidation.

CHAPTER XXVIII. Foods. — The foods are organic compounds which obey the same laws as other chemical compounds. They contain carbon, hydrogen, oxygen, and nitrogen, phosphorus, sulphur, chlorine, iodine, sodium, potassium, calcium, magnesium, and iron. These elements are found in the structure of the body, and the foods to supply the body structure must be composed of these elements. Heat and power for the body are obtained by the oxidation of the carbon and hydrogen of the foods. The five simple classes of nutrients into which the foods are divided are the proteids, carbohydrates, fats, inorganic salts, and water. The carbohydrates and fats are made up of carbon, hydrogen, and oxygen. When their carbon and hydrogen burn, carbon dioxide and hydrogen oxide (water) are formed.

Proteids contain nitrogen in addition to carbon, hydrogen, and oxygen. They sometimes contain sulphur. The proteid is the most necessary member of the body-building nutrients. No other food can take its place. Nitrogen is of absolute necessity to the body and the proteids are the nitrogenous foods. The inorganic salts give the mineral elements which build up the body skeleton. The water supplies the necessary water of the body which is constantly being lost.

Soaps are made by the neutralization of fatty acids (existing in fatty compounds similar to those of the food) with hydroxides like those of potassium and sodium. When soap fat is boiled with an alkali, glycerine and fat separate off and the acids neutralize, forming sodium salts. When the soap is dissolved in water some of the caustic soda which was originally present re-forms. This alkali cleanses off the oily matter from the skin, and the other part of the soap, together with the water, takes off the dirt mechanically.

CHAPTER XXIX. Clothes. — Cloth is made from woven vegetable and animal fibers. Vegetable fibers are plant cells of simple structure. Animal fibers are much like proteid in composition and are generally complex cell structures or bundles of cells. Sheep give wool and the silkworm gives silk fibers. These are animal fibers. The cotton and flax plants furnish vegetable fibers. Long, coarse wool makes worsted goods and very fine wool makes woolen goods. Both wool and silk commonly contain carbon, hydrogen, oxygen, nitrogen, and sulphur. Wool has to be thoroughly cleansed and bleached. Cotton is bleached and boiled with soda ash for cleansing. Linen bleaching is carried on out of doors.

The first aniline dye (dye made from coal tar) was discovered in 1856 by Sir William Henry Perkin. The great majority of the dyes used to-day are coal-tar compounds. Dyeing depends on the deposit of some insoluble substance within the fiber. Animal fibers can be dyed directly. Vegetable dyes are dyed indirectly. In indirect dyeing, two chemicals, a dye and a mordant, precipitate the color desired within the fiber. The substance formed by the interaction of a dye and mordant is called a lake. Alizarine illustrates the use of a mordant and a dye. The reaction between these two is generally a type of acid + base = salt + water. Alizarine and indigo are two dyes which are synthetically made in the laboratory which have revolutionized agricultural conditions in two great territories.

CHAPTER XXX. HEAT AND LIGHT IN CHEMISTRY. — There are two types of chemical reaction, the endothermic or heat-absorbing and the exothermic or heat-developing. These reactions are measured by a unit of heat, the calory, which is the amount of heat required to raise one gram of water one degree Centigrade. The specific heat of any substance is the amount of heat required to raise one gram of any substance one degree Centigrade in temperature. The synthesis of water is an exothermic reaction which liberates 68,400 calories in forming eighteen grams of water at 0° C. The analysis of water is an endothermic reaction which requires 68,400 calories to decompose eighteen grams of water at 0° C. into hydrogen and oxygen. The law for this action is that the amount of energy manifested in the reaction is equal to that manifested in the reverse reaction.

Coal gas is produced by the destructive distillation of coal which produces coal tar, ammoniacal liquor, and coal gas.

Combustion occurs when the temperature of a substance is raised above the ignition or kindling point. In ordinary exothermic combustion the ignition point of a substance is lower than the temperature produced by this burning. Flame occurs only when both the substances combining are in the form of gases or vapors. The construction of a flame requires an outer cone and a hollow inner part. The light of the flame may be due to the solid matter present in the flame, to the density of the flame gas or to dilution and cooling. There are two Bunsen burner flames, the outer and inner flame, the oxidizing and reducing. The very high heat of the electric furnace is reached by the passage of electricity from one terminal to another in a fashion very similar to the passage of current in the arc light. The heat evolved during this passage is commonly above 3500 degrees C., and by the construction of the furnace this heat is made to act directly upon a crucible.

Light reactions in chemistry reach their application in photography, where light acts upon silver salts, breaking them down, the decomposition varying with the amount of light thrown on the plate. Positive prints are made which are the opposite in light and shade of the negative and the development of the photograph depends upon the use of some reducing agent such as pyrogallic acid or ferrous sulphate.

CHAPTER XXXI. IRON AND COPPER. — Iron-working depends primarily upon the reduction of iron ore to metallic iron by heating the ore with carbon. Iron owes its great value to its peculiar metallic properties, its tenacity, its lightness, its rigidity and its infusibility. It is divided into three classes, wrought iron, cast iron, and steel. The production of these three classes is dependent on the amount of carbon present in the iron. Wrought iron, which must be forged and welded, contains the smallest quantity of carbon. Steel, which is fusible and malleable and which can be so tempered as to be hard and brittle or soft and tough, contains more carbon than wrought iron and less than cast iron. The tempering of steel depends on the swiftness with which it is cooled from its heated state.

Iron is reduced from ores, which are generally oxides, by carbon. The iron sulphides are seldom used for supplying metal. Magnetite, or loadstone, is natural magnetic iron. The middle West of the United States and part of the South contain a large portion of the iron which can be readily worked. Iron is, however, found in varying quantities throughout the United States. Three raw materials are needed in iron production, iron, coal, and limestone. Limestone is used for separating ore from the clay with which it is mixed, forming a glass with the clay through which the metal can fall. The carbon of the fuel removes oxygen from the iron and forms carbon dioxide. Iron is prepared from its ores by heating coal, limestone, and iron in a blast furnace, through which great currents of air are drawn. A common test for iron is the production of iron hydroxide precipitated when iron compounds are treated with an alkaline hydroxide.

Copper. — Copper is the best conductor of electricity that can be obtained at a moderate price. The metal is hard and tough and is easily drawn out into wires. Ordinary copper may be oxidized to black copper oxide, turning the flame green as it changes. A common test for copper is the precipitation of copper sulphide by means of hydrogen sulphide. Copper is used for bolts, rivets, currency, electrotypes, and for vessels of various descriptions. Copper forms several valuable alloys, brass, bronze, and nickel alloy especially. The compound copper sulphate is of value in dyeing. The ores of this metal are found largely on the shores of the Great Lakes, in Montana and in Arizona. It is chiefly reduced from the sulphide, carbonate and oxide ores. The first two are roasted to change them to oxide. The latter is reduced by roasting with coke. Crude copper may be purified by means of an electrolytic separation.

CHAPTER XXXII. SODIUM, POTASSIUM, CALCIUM, AND ALUMINUM. — These four elements are metals of the soil. Sodium colors a flame yellow. This action is a type of a flame test for the metals. Each metal when incandescent imparts a peculiar color of its own to a flame. Metallic sodium was first made by Sir Humphrey Davy through the electrolytic separation of sodium hydroxide. To-day sodium is produced by the electrolysis of fused sodium hydroxide. This element is a light silver-colored solid which eagerly unites with water through the evolution of hydrogen and forms sodium hydroxide. It must be kept under oil. Sodium biborate is borax and is made up of sodium, boron, and oxygen. Borax is largely used in galvanized tin and in enameling. It is used in welding metals and in cleansing. Metallic potassium is obtained by an electrolytic separation similar to that used in preparing sodium. Potassium salts are chiefly found in natural deposits similar to those at Stassfurt. This element, like sodium, oxidizes swiftly in the air, produces hydrogen and potassium hydroxide by its reaction with water, and must be kept under oil. Its flame test is violet. Its compounds are used for gunpowder, for making oxygen and for making soap.

Calcium is the metallic part of lime, used widely for building. Lime, the oxide of calcium, is an alkali used almost as widely as sulphuric acid. 2.5 per cent of the earth's crust is sodium, about the same percentage is potassium, 3.5 per cent is calcium, 8 per cent is aluminum, 50 per cent is oxygen, around 25 per cent is silicon. The element calcium is golden yellow in color and is comparatively hard and malleable. Calcium is formed by the reaction between calcium iodide and sodium, which produces calcium and sodium iodide. Aluminum is prepared by an electrolytic separation, similar to that used in the separation of sodium and potassium. Its great lightness, combined with its uncommon strength, makes the metal of great value. The ruby, sapphire, topaz, amethyst, emerald, turquoise and garnet, are all compounds of aluminum. Corundum is also a compound of aluminum.

Gold and Silver. — Gold and silver are largely of value for coinage and jewelry. Gold is found in the United States, in South Africa, Australia, and Russia. If this element is found in sand it can be separated by washing, a process which depends on the differing density of the heavy gold and the light sand. If found in quartz the ore must be crushed and dissolved as a cyanide. The metal is afterward separated out by the electric current. Gold is itself soft, but is rendered hard by the metals which are alloyed with it for common use. The carats which measure gold are figured in parts per 24. Pure gold is 24 carats.

Silver, like gold, does not oxidize at ordinary temperatures in the air. It tarnishes when exposed to the action of sulphur. It is extracted from its ores by mercury amalgams. The mercury unites with silver in the mass of crushed ore to form an alloy. When heated, the mercury of the amalgam goes off as a vapor and leaves the silver behind. This metal is a great conductor of heat and electricity and its salts break down under the influence of light.

CHAPTER XXXIII. Lead, Zinc, Tin, Arsenic, and Antimony. —

Lead is of great value because of its softness. It is also valuable because of its fusibility. It serves many purposes, such as the making of lead shot, solder, and type metal, the lining of chambers for sulphuric acid, and making lead pipes. It swiftly oxidizes in the air. It is readily reduced from its ore galena by reduction when roasted. It is then heated out of contact with the air. Taken internally this element is poisonous, and should not be used for pipes in a water system when the water contains sulphates or carbonates.

Zinc is a light metal with considerable rigidity and is of value for gutters and conductors. It is very fusible and readily crystallizes. It forms valuable alloys with copper. It is reduced from its ores by oxidation and reduction.

Tin is of value in coating iron to keep it from the action of chemicals which would injure it. With lead it forms solder and pewter. It is so soft in its purely elementary form that it can be cut with a knife, it melts at a low temperature and is reduced from its ores by roasting, washing, and smelting.

Arsenic and antimony stand on the border line between the metals and non-metals, and are sometimes called metalloids. Arsenic looks like the metals and in some ways acts like them. Its compounds are more like the non-metals. With hydrogen it forms arsine, a very poisonous gas. This gas is generated at ordinary temperatures when organic matter and arsenic are present together. Consequently wall paper or hangings which contain arsenic should never be used in living-rooms. Arsenic is prepared by roasting its ores and condensing the vapor which comes off.

The alloys of antimony make this metal of special value. This is due largely to the property which antimony alloys possess of expanding when they change from a liquid to a solid state. This property gives its sharp lines to type-metal which is made up of lead, antimony, and tin. Antimony burns, forming antimony trioxide, and with hydrogen it forms stibine (antimony hydride). It is reduced from its ores by heating with scrap iron.

PART III

QUESTIONS

CHAPTER I

1. How can we think of the universe as built up from the stars which we see at night?

2. Why is it we can see the units which we may consider as composing the universe, and not the units which compose a rock or a tree?

3. What is the difference between hypothesis and law?

4. Illustrate a hypothesis by stating our belief that there are stars we have not seen.

5. What is the hypothesis commonly held in regard to the constitution of matter?

6. State the atomic hypothesis.

7. Define a molecule.

8. How is the total mass of any body built up?

9. Define an atom.

10. What is the relation between the atom and the molecule?

11. What is the belief concerning the size of a molecule?

12. In what peculiar way do atoms act in chemical actions?

13. State a result of the indivisibility of the atoms as regards a bar of iron.

14. How many different atoms are there so far as we know?

15. How many simple molecules or masses are there?

16. What is an element?

17. Can we have elementary and compound atoms?

18. Can we have elementary and compound molecules?

19. How would you form an elementary molecule?

20. How would you form a compound molecule?

21. About how many elements are in common use in the formation of compounds?

22. Give some of the elements that you know.

23. Give some of the compounds that you know.

24. What are the three states of matter?

25. How does water show the three states of matter?

26. Name some other substances which are found in all the three states.

27. What two things commonly determine the state of matter in which any substance may be?

28. What elements are gases? What elements are solids? What elements are liquids?

29. What metals do you know?

30. What non-metals do you know?

31. What are the chief differences between metals and non-metals?

32. Define a physical change.

33. Define a chemical change.

34. State some physical change which you have known.

35. State some chemical change which you have known.

CHAPTER II

Note. — In connection with this chapter examine the tables on the back cover page.

1. What means have we of testing the substances that we meet?

2. How could we state the properties of a piece of chalk in terms of the five senses?

3. What test replaces touch and is far more accurate?

4. What is the common reason for a body's change in weight?

5. What exceptions are there to this general statement?

6. What does loss in weight commonly mean with regard to change in structure?

7. What does gain in weight commonly mean with regard to change in structure?

8. Give three heads under which chemical change commonly takes place.

9. Define analysis, synthesis, and metathesis.

10. How do we use the sense of touch as a balance in daily life?

11. State some uses of weight comparisons in every-day life.

12. What are the two great measuring systems of the world?

13. What is the chief difficulty with the English system?

14. What is the chief advantage of the metric system?

15. What system, similar to the metric system, do we have in the United States?

16. Give the metric scale of length.

17. How do you write a length of two centimeters, four millimeters, numerically?

18. How long is a meter in English inches?

19. How long is a kilometer in miles?

20. What is the square system of metric measurement?
21. What is the cubical system of metric measurement?
22. Give the divisions of the metric system of weights.
23. How would you write four grams, six decigrams, five centigrams, two milligrams, numerically?
24. How would you write nine grams and sixteen milligrams numerically?
25. How many cubic centimeters in a liter?
26. What part of a liquid ounce is a cubic centimeter?
27. What are the two thermometer scales?
28. State the differences between the two.
29. Why is the Centigrade scale used in scientific work?
30. What is density?
31. Is the density of a body that will float in water more or less than 1? Is the density of a body that will sink in water more or less than 1?
32. What would be the density of a piece of metal which weighed forty grams and had a volume of five cc.?
33. What is the density of water in the metric system and in the English system?
34. If I have forty cubic centimeters of water, how many grams have I?
35. Give some practical uses of density.
36. How can we use gaseous densities in collecting gases?
37. How are gases collected over water?
38. How are gases collected by upward displacement?
39. How are gases collected by downward displacement?
40. Give some of the means by which such chemical changes as the production of gases are brought about.

CHAPTER III

Oxygen

1. What gas is essential for burning?
2. On what does the heat and flame of ordinary burning depend?
3. Why do we open the drafts of the stove to increase the fire?
4. Why do we throw water or a blanket over the flame to put out fire?
5. How do we use oxygen in the human body?
6. Where do we get the oxygen we use for fire?
7. What are the proportions of oxygen and nitrogen in the air?
8. What is the use of nitrogen in the air?
9. What is the use of oxygen in the air?
10. What general action happens in the stove when a hodful of coal is burned?

11. What happens to air as it passes through a fire?

12. What goes up the chimney from a fire below?

13. What is the difference between the increase in weight of a metal when it is burned and the apparent decrease in weight of wood when it is burned?

14. What is added to the burning metal?

15. What happens to the burning wood of a match?

16. Who discovered oxygen?

17. How did he do it?

18. When red oxide of mercury is heated in a tube, what goes off? What is left behind?

19. What happens when a glowing splinter is thrust into the gas produced by heating red oxide of mercury?

20. How can we define burning?

21. Why is burning swifter in oxygen than in ordinary air?

22. What is the common laboratory method of preparing oxygen?

23. What is the use of potassium chlorate in the mixture?

24. What is the use of manganese dioxide?

25. What is a catalytic agent?

26. How can we collect oxygen?

27. What are the properties of oxygen? Its solubility? Weight? Color? Odor?

28. Why do the bubbles of gas force water out of the bottle?

29. What is the result of plunging burning substances into oxygen?

30. What is the test for oxygen?

31. What are the compounds of oxygen?

32. Where are these compounds found in nature?

33. In what form are metals often found in mines?

34. How can we prepare oxides?

35. Name the compounds formed when iron, copper, magnesium, and carbon burn.

36. How many elements combine directly with oxygen?

37. What is the special value of carbon's oxidation?

38. Compare swift and slow oxidation in the burning of a match and the rust of iron.

39. Give some examples of oxidation in nature.

40. Explain oxidation in the body.

41. What is the use of oxides in fireproof materials?

42. Does oxygen commonly combine with substances at the temperature of the air?

43. How is a fire lighted by a match?

CHAPTER IV

HYDROGEN

1. What did Henry Cavendish do, and what was his especial gift to learning?

2. What great fact concerning hydrogen and its action did he discover?

3. How is water formed? What is water, chemically speaking?

4. What is the chemical standard of atomic weights?

5. Why was hydrogen chosen for this standard?

6. Where is hydrogen found in nature?

7. How widely is it found in combination?

8. Why does a balloon rise in air?

9. What does this show as regards the comparative weight of gases?

10. How can we use a gas, especially hydrogen, for measuring?

11. What is the weight of a liter of hydrogen at normal temperature and pressure?

12. How much heavier would a gas be than hydrogen if it weighed .45 grams?

13. What is the weight of a liter of air if air is 14.4 times as heavy as hydrogen?

14. From the analysis of what group of substances in which hydrogen is always present do we obtain it?

15. What is the easiest way of getting hydrogen from this class of bodies?

16. What is the simplest acid and metal to use in obtaining hydrogen?

17. Explain the process of preparing hydrogen.

18. Give the properties of hydrogen, color, odor, weight, and solubility.

19. In what ways do we collect hydrogen?

20. What is left in the generator after the action?

21. Explain carefully just what goes on in the generator when the zinc displaces the hydrogen of the acid.

22. Write the equation for forming the gas.

23. Does hydrogen support combustion?

24. What is the test for hydrogen?

25. What precaution must be taken before lighting hydrogen?

26. What is formed when hydrogen burns in the air?

27. How can the heat of the hydrogen flame be used?

28. What are the characteristics of hydrogen dioxide, the second oxide of hydrogen?

29. How is it used?

30. How can it be formed?

31. How can we get hydrogen from steam and red-hot iron?

32. What is the value of hydrogen in the air-ship?

CHAPTER V

Acid, Base, and Salt

1. What are three examples of acid, base, and salt which are found in every kitchen?

2. Where do we find the acids in nature?

3. Give the names of half a dozen acids known to you.

4. What are the three great acids of the commercial world?

5. Give the common and the chemical names of these great acids.

6. What part does hydrogen play in making acids?

7. Name the three great acids as hydrogen compounds.

8. What other element, beside hydrogen, is commonly present in acids?

9. What class of substances must be present in acids?

10. What characteristic of acids does the making of gingerbread exemplify?

11. Explain what happens when a seidlitz powder is mixed.

12. What occurs when a carbonate meets an acid?

13. What can you say concerning the destroying properties of acids?

14. What effect does an acid have upon litmus?

15. What is litmus?

16. What is the action of acids on methyl orange?

17. What is methyl orange?

18. What is an indicator?

19. Give the general characteristics of acids.

20. What happens when a bee stings?

21. Why do we put ammonia on a bee sting?

22. Why do we add limewater to neutralize sour milk?

23. What is the general action when an acid meets a base?

24. What is the action of a base on cloth?

25. What is the action of a base on litmus, and methyl orange?

26. What are the two indicators which belong especially to bases?

27. What is the action of bases on turmeric and phenolphthalein?

28. Where do turmeric and phenolphthalein come from?

29. How is sodium hydroxide prepared when we drop sodium into water?

30. Explain the action which took place when the sodium hydroxide was formed.

31. What is hydroxyl?

32. How do metals cause a difference between acids and bases?

33. Give the definition of a base.

34. What is the effect of adding hydrochloric acid to sodium hydroxide?

35. Explain the action which occurs.
36. Of what general rule is this action an example?
37. Give the equation which takes place under such circumstances.
38. How are salts produced?
39. Give some substances in which the normal salts are formed.
40. What is the difference between a normal and an acid salt?
41. What is a basic salt?

CHAPTER VI

WATER

1. Give the general distribution of water over the face of the earth.
2. Give some examples of the solvent power of water.
3. Explain the properties of a physical solution such as a water solution.
4. How do the dissolving and suspending powers of water change the face of the earth?
5. What is the action of water on the soil?
6. How does the rain give food to plant and animal life?
7. How much of the body's weight is water?
8. Give some of the percentages of water in our daily food.
9. Give some of the uses of water.
10. Why can we not get pure water in nature?
11. What happens when the rain strikes the earth?
12. Why is it dangerous to drink unclean water?
13. What precautions should be taken as regards drinking water?
14. From what does the danger of water-borne diseases really come?
15. What kinds of soils purify water?
16. What kinds of soils are dangerously likely to leave water unpurified?
17. What is the chief action of water on soap?
18. How does boiling affect some hard waters?
19. What is the "brown fur" of the kettle?
20. Why does hard water containing carbonates become soft when boiled?
21. When is hard water permanently hard?
22. What is the effect of hard water on soap?
23. What is the danger of hard water in a steam boiler?
24. Explain the action of heat on a boiler which has become incrusted?
25. What are stalactites and stalagmites and how are they formed?
26. What is water of crystallization?
27. Explain its action on washing soda.
28. Define efflorescence and deliquescence. Illustrate them.

29. How do lithia waters become charged?

30. How can we make practically pure water in the laboratory?

31. Describe the process of distillation.

32. How can we prove by analysis that water is made up of two parts hydrogen to one of oxygen?

33. How can we prove by synthesis that water is made up of two parts hydrogen to one of oxygen?

34. How does the expansive property of steam act in the steam engine?

35. Explain the general action of the steam engine.

36. What is the general rule of expansion and contraction when substances are heated or chilled?

37. How does water act when heated or chilled?

38. What is the effect of the unusual expansion of water?

39. Explain the action which takes place in a lake cooling on a fall day.

40. Why does not the ice freeze solidly from the surface of such a lake to the bottom?

41. How can fish live in winter waters?

42. What would happen if the anomalous expansion of water ceased?

43. How can we say that on the anomalous expansion of water depends life in the temperate zones?

CHAPTER VII

LAWS

1. What difference is there between the laws of man and those of nature?

2. Why are the laws of man not more infallible than they are?

3. How are nature's laws demonstrated?

4. What is the necessity for a knowledge of nature's laws?

5. How are the great underlying laws of the universe proved?

6. What is the law of the conservation of energy?

7. Illustrate this law by the action of coal burned in the furnace.

8. What is the equation which shows the relation between total energy, useful energy, and waste energy?

9. Why is perpetual motion impossible?

10. How is the law of the conservation of energy applied in the human body?

11. What effect does this law have upon our food?

12. State the law of the conservation of matter.

13. Illustrate this law by the combination of carbon and oxygen.

14. What is the relation of this law to the total matter of the universe?

15. What effect does time, even centuries of time, have upon this law?

16. State the law of definite proportions by weight.

17. Illustrate this law by the salt of the sea.

18. Illustrate this law by common salt prepared in nature and in the laboratory.

19. Illustrate this law by sand.

20. State the law of multiple proportions by weight in its first form.

21. What is a constant weight?

22. State the law of multiple proportions in its second form.

23. Give the illustration of the stable filled with horses.

24. Explain the difference between a simple ratio and a fractional ratio. Illustrate this law by the oxides of hydrogen. Illustrate it by the oxides of carbon and by the oxides of nitrogen.

25. Explain what simple multiples are.

26. State the law of reciprocal proportion both in the first way and in the second.

27. Illustrate this law by the comparison of white and bay horses.

28. Show how sulphur and oxygen in their combination with hydrogen illustrate this law. Do the same with sulphur and oxygen in combination with sodium.

29. Who was the man who expounded the atomic theory?

30. According to this theory, how does chemical combination take place?

31. How are atomic weights obtained?

32. What is the weight of one atom of hydrogen called?

33. How great is the atomic weight of one atom of oxygen?

34. What is the weight of an atom of sulphur?

35. What is the effect of the indivisibility of the atom on the law of constant proportions?

36. What is the effect of the same thing on the law of multiple proportions?

37. How does the atomic theory agree with the law of reciprocal proportions?

38. What relation do these laws have to every-day manufacture?

39. What would be the effect of carrying on manufacturing processes without a knowledge of these laws?

CHAPTER VIII

CHLORINE

1. What is the gaseous element which bleaches?

2. Where is chlorine found?

3. Name some of the compounds which contain chlorine.

4. How is hydrochloric acid made?

5. How can we get chlorine from hydrochloric acid by the electric current?

6. What is the density of chlorine?

7. What are its properties? What, especially, is its solubility?

8. How can we make chlorine in the laboratory?

9. How does manganese dioxide react on hydrochloric acid to form chlorine?

10. How does chlorine act toward a flame?

11. How does it act toward dry litmus paper?

12. How does it act toward wet litmus paper and wet colored cloths?

13. What happens when chlorine and hydrogen are mixed in the light?

14. What is the action of chlorine on steam?

15. What is the general part that oxygen plays in bleaching with chlorine?

16. How does this explain the fact that wet cloths bleach with chlorine when dry cloth will not bleach?

17. What is the action of light on chlorine water?

18. What kind of goods are especially likely to be bleached with chlorine?

19. What is the action of this gas on plant life?

20. What is its action on the human frame?

21. How can chlorine be stored for commercial use?

22. What is bleaching powder and how is it made?

23. How is bleaching powder used for bleaching?

24. What is the common name for hydrochloric acid?

25. Where is this acid found in nature?

26. From what substance is it prepared in the laboratory?

27. What occurs when salt and sulphuric acid are heated strongly?

28. What action generally goes on in the laboratory process for making hydrochloric acid?

29. What are the properties of the gas, as regards color, weight, and solubility?

30. What gases have we studied which are insoluble in water?

31. How does hydrochloric acid gas compare with these gases in solubility?

32. What is the manufacturing process for making hydrochloric acid?

33. In what process is hydrochloric acid a by-product?

34. What is the test for hydrochloric acid?

35. Give the reaction which occurs when we add hydrogen chloride to silver nitrate?

36. What is the action of the gas on wet and dry litmus?

37. What is the chemical composition of the laboratory solution of hydrochloric acid?

CHAPTER IX

The Chlorides

1. What salts are found in the ocean?
2. Does the ocean contain different amounts of solids?
3. How did our ancestors obtain their salt?
4. How is essentially this same process used to-day?
5. How much (by English weight) does a gallon of sea water contain of solid matter?
6. What are salt licks?
7. Where is most of our American salt manufactured?
8. Whence does most of the foreign salt come?
9. How can we tell when the salt of our table is not absolutely pure?
10. What compounds present in table salt attract moisture?
11. Name four groups formed from four acids.
12. How are chlorides produced?
13. What is deliquescence?
14. What is the chief characteristic of calcium chloride?
15. How is calcium chloride used in drying gases, or protecting balances from rust?
16. What is the chief property of cobalt chloride?
17. How does the solution of cobalt chloride act as a sympathetic ink?
18. How is cobalt chloride used in toy barometers?
19. What is calomel?
20. What is corrosive sublimate?
21. Compare these two chlorides of mercury.
22. What is the danger of using corrosive sublimate?
23. What are the chief uses of potassium and magnesium chlorides?
24. Where are these two chlorides found?
25. What peculiar property do the chlorides of gold and silver possess?
26. How is tin chloride used?
27. What is the use of zinc chloride?
28. What is the danger of a careless use of this salt?
29. What is the use of barium chloride?
30. What is the use of copper chloride?
31. What is the use of iron perchloride?
32. What is the chief method of making chlorides?
33. Name four metallic chlorides.
34. What is the general solubility of the chlorides?

35. Give two reactions for preparing zinc chloride.

36. Give the reaction when hydrogen is added to the hydroxide of the metal.

37. Of what general reaction is this an example?

38. How can chlorides be made without the use of hydrochloric acid?

39. Give the reaction for making sodium chloride from sodium and chlorine.

40. What is the test for the chlorides?

41. What three metals form insoluble chlorides with hydrochloric acid?

42. Name the compounds formed.

CHAPTER X

Fluorine, Bromine, and Iodine

1. What are the four members of the halogen family?
2. Why was this family called the halogens?
3. Is this the only family group in chemistry?
4. Why is it particularly studied as a family?
5. Which is the most active of the halogens?
6. How is glass etched by fluorine?
7. What can you say about the combining powers of fluorine?
8. What is the commonest compound of fluorine?
9. Describe fluorspar.
10. How is hydrofluoric acid made from fluorspar?
11. Give the properties of hydrofluoric acid.
12. Explain the process of etching by the use of hydrofluoric acid.
13. How is elementary fluorine obtained?
14. What chemist accomplished this separation?
15. How was it done?
16. Give the properties of fluorine.
17. By what test can we recognize fluorine and hydrofluoric acid?
18. Where does bromine occur in nature?
19. How does bromine differ from chlorine and fluorine?
20. How is it chiefly characterized?
21. Give the properties of bromine.
22. How is bromine made in the laboratory?
23. What is the acid which bromine forms with hydrogen?
24. What is the test for bromine?
25. How is bromine used commercially?
26. Give the relative weights (atomic weights) of the halogens.
27. Where is iodine found?
28. Tell the story of its discovery by Courtois.

29. Give the properties of iodine.

30. What is sublimation?

31. What is meant when we say that iodine sublimates?

32. Give the starch test for iodine.

33. What compounds do these four halogens form with metals?

34. What do we find concerning the relative energy of these four elements?

35. What four acids are formed by these four elements?

36. Describe the formation of these acids.

37. What is true of the replacing power which one of these elements possesses in regard to another?

38. What law places the elements in the family groups where they belong?

39. Is there any element which seems to have no relation to any other element?

CHAPTER XI

CARBON DIOXIDE

1. How is carbon dioxide formed in nature?

2. Where is it found in nature?

3. What is the reason for its great production?

4. How is most of the carbon dioxide which is used commercially prepared?

5. Give the reaction for this process.

6. What is the laboratory method for making carbon dioxide by the use of hydrochloric acid?

7. Give the reaction for this method.

8. Why cannot sulphuric acid be used for the process as well as hydrochloric?

9. Give the reaction which goes on when sulphuric acid is added to calcium carbonate.

10. Give the characteristics of carbon dioxide gas.

11. What can you say of the production of carbon dioxide in close or crowded rooms?

12. Why does not the air become impregnated with carbon dioxide?

13. Explain the "cycle of carbon."

14. What is the chief use of carbon dioxide in nature?

15. What is the soda of the soda fountain?

16. How is it made and kept?

17. How do fire extinguishers use carbon dioxide?

18. What reaction goes on when an ordinary fire extinguisher is used?

19. What is the test for carbon dioxide?

20. Give the reaction for this test.
21. What happens when bread is raised with yeast?
22. What occurs when soda biscuits are made?
23. Where is carbon monoxide found in an open fire?
24. Describe the action of the gases in an open fire.
25. What is the use of carbon monoxide in iron-working?
26. How is water gas made?
27. Give the reaction for its formation.
28. What is added to water gas to give greater light?
29. Why is water gas not so satisfactory as coal gas?
30. Explain the formation of carbon monoxide in the laboratory.
31. Give the reaction for its preparation.
32. What is the test for carbon monoxide?
33. Give the reaction for this test.

CHAPTER XII

THE CARBONATES

1. What is the especial use of calcium carbonate?
2. What are the common names of the various kinds of calcium carbonate?
3. Where is calcium carbonate found in nature?
4. What is the most valuable form in which calcium carbonate appears and why is it the most valuable?
5. Where is sodium carbonate found in nature?
6. What is the common name for sodium carbonate?
7. What is the common name for sodium bicarbonate?
8. What is the one chief use of such lakes as the Great Salt Lake of Utah?
9. Where is potassium carbonate found in nature?
10. What is the common name of potassium carbonate?
11. What is lye?
12. In making what household necessity is it used?
13. What is the test for a carbonate?
14. Give the reaction for the test.
15. How can carbonates be made in the laboratory?
16. What are the two manufacturing processes by which sodium carbonate is made?
17. Where does the calcium carbonate used in commerce come from?
18. How is potassium carbonate prepared?
19. Tell the story of the way in which the Le Blanc process came to exist.
20. Explain the Le Blanc process.

21. Give the reactions involved in the Le Blanc process.

22. What are the advantages of the Le Blanc process?

23. Explain the Solvay process.

24. Give the reactions for the Solvay process.

25. What are the characteristics of the carbonates?

26. What effect does the insolubility of the carbonates have upon their usefulness for building?

27. What is carbonic acid?

28. How is it prepared?

29. What are the uses of the carbonates in nature?

30. Explain the action of mortar.

31. What is meant by the "setting of mortar."

32. What is the action of limestone in iron working?

33. Explain the making of glass.

34. What kind of glass is used for eyeglasses?

35. What differences exist between making such glass and common glass?

36. What are the valuable qualities of glass?

37. What is chalk?

38. On which of its properties does the schoolroom use of chalk depend?

CHAPTER XIII

EQUATIONS

1. What is chemical affinity?

2. What is the symbol of an element?

3. For what do O and H stand?

4. How could George Jones be used for the symbol of a man?

5. What is a microcrith?

6. What is the symbol for water?

7. What does H_2O signify?

8. What does SO_2 signify?

9. Give the symbol for nitrogen, carbon, and platinum.

10. Where does the symbol of salt come from? of gold, tin, iron, mercury, and water?

11. How do the names of the more recently named symbols differ from the elements named in ancient times?

12. What does the symbol HCl indicate?

13. Write the reaction which occurs when calcium carbonate is heated?

14. Under what circumstances is a reaction reversible?

15. What is meant by writing arrows instead of the sign of equality in reactions?

16. What are the three classes of equations?

17. What two groups of phenomena must be considered in the study of reactions?

18. What are the differences between these?

19. How can chemical affinity be roughly measured?

20. How do reactions between substances in solution differ from other reactions?

21. What happens when sodium chloride is dissolved in water?

22. Explain the action which occurs if we have sodium chloride in solution and pour in silver nitrate.

23. Write the reaction which takes place, using full names.

24. Give the illustration of the two fleets.

25. Write the reaction in symbols.

26. What change occurs to silver chloride when it comes out of solution?

27. What are electrolytes?

28. What is the law of attraction and repulsion of bodies charged with like and unlike electricity?

29. What group of substances are regularly electro-positive?

30. What group of substances are regularly electro-negative?

31. What is valence?

32. What is a monovalent substance?

33. What is a bivalent substance?

34. What is a trivalent substance?

35. Illustrate this by the two chlorides of iron and the chloride of sodium.

36. Write reactions for every symbol of the non-metallic elements and radicals of the list combined with every symbol of a metal. For example:

$AgCl$
$AgNO_3$
Ag_2O
Ag_2SO_4

CHAPTER XIV

Nitrogen and Ammonia

1. What is the special use of nitrogen?

2. What would happen if nitrogen were not so great a part of the air?

3. What can you say about the combination of nitrogen with other substances?

4. Where is nitrogen found in nature?

5. What part of our food is nitrogenous?

6. What does nitrogenous mean?
7. Give the method of separating nitrogen from the oxygen of the air.
8. Give the two other laboratory methods for making nitrogen.
9. Give the reactions for these methods.
10. Why was ammonia called alkaline air?
11. Who separated out and studied ammonia gas?
12. Where is ammonia found in nature?
13. What does the presence of ammonia in bodies of water denote?
14. Why will ammonia feed the plants when the nitrogen of the air will not?
15. What is the use of ammonia in fertilizers?
16. In what waters does ammonia occur naturally?
17. How can ammonia gas be prepared in the laboratory?
18. Where does most of our ammonia come from?
19. What are the tests for ammonia?
20. How could you determine the presence of ammonia in a compound?
21. What are the commercial uses of ammonia?
22. What is the principle on which the cooling action of ammonia depends?
23. Describe the changes which take place in liquefying ammonia gas.
24. How is artificial ice made?
25. How do cold storage plants work?
26. What is the difference between ammonia gas and ammonia water?
27. Why is ammonium called the theoretical metal?

Questions on the Oxides of Nitrogen

1. Give the five oxides of nitrogen with their symbols.
2. Give the law which they illustrate.
3. How is nitrous oxide made? Give the reaction for its preparation.
4. What are the characteristics of nitrous oxide?
5. What is its action toward glowing wood?
6. How may it be differentiated from oxygen?
7. What is its use in medicine and dentistry?
8. What is its common name?
9. Why does it have that name?
10. How is nitric oxide made?
11. Give the reaction for its preparation.
12. What are the properties of nitric oxide?
13. In what reactions does the gas commonly appear?
14. How is nitrogen trioxide made?
15. What is its appearance?
16. How is nitrogen peroxide made?

17. What is its appearance?
18. In what different forms does it appear?
19. What are the characteristics of nitrogen peroxide?
20. How is nitrogen pentoxide made?
21. Give the reaction for both the methods of making it.
22. What are the characteristics of nitrogen pentoxide?

CHAPTER XV

Nitric Acid and the Nitrates

1. What is the special use for nitric acid?
2. Where does it occur in nature?
3. What are its salts?
4. Where are they obtained in nature?
5. Describe the process of making nitrates for fertilizer from the nitrogen of the air.
6. Where in nature does this same process occur?
7. Give the reaction for this preparation.
8. Describe the common process for making nitric acid.
9. Give the reaction for its manufacture when slightly heated; when strongly heated.
10. Give the characteristic properties of nitric acid.
11. What is the difference between the action of nitric, hydrochloric, and sulphuric acids on cloth?
12. What is one essential difference between the action of nitric acid and other acids on metals?
13. Define and illustrate oxidation.
14. Define and illustrate reduction.
15 Describe the copper nitrate test for nitric acid.
16. Give the reaction for this test.
17. Describe the brown ring (ferrous sulphate) test for nitric acid.
18. What is the use of nitric acid in testing for gold?
19. Explain the use of the touchstone and the touch-needles.
20. What is "aqua regia?"
21. What is the action of "aqua regia" on gold? What compound of gold is produced by this liquid?
22. What is the relation of nitrates to explosives?
23. Give some of the uses of explosives.
24. Of what is gunpowder composed?
25. What happens when gunpowder or other explosives explode?
26. What happens when a bullet is shot from a gun?
27. What kind of plants need nitrogen most?
28. How do they get their nitrogenous food?

29. How does a modern sewage plant act?
30. What is the action of nitrifying bacteria on sewage?
31. Give the formula for nitrous acid.
32. Where are nitrites found in nature?
33. How are nitrites commonly formed?
34. What is the test for the presence of nitrites?
35. Why is the presence of nitrites in drinking water considered a sign of danger?

CHAPTER XVI

AIR

1. Why can we say that we live at the bottom of an ocean of air?
2. How deep is the ocean of air?
3. Is all this ocean of the same density?
4. Where is the air rarest?
5. Where is it most dense?
6. What changes in the density of the air do you find in going up a mountain?
7. Explain this variation of density on different levels by the illustration of the haymow.
8. Why does air change its density on the same level?
9. What particular use can be made of the compressibility of air?
10. Why does air vary so greatly in density?
11. What is the approximate proportion of oxygen and nitrogen in air?
12. How can we prove by the use of phosphorus that air is about four fifths nitrogen and one fifth oxygen?
13. Why in this experiment does water rise and take the place of phosphorus oxide?
14. What other substances exist in air beside oxygen and nitrogen?
15. When was argon discovered?
16. Give the names of some of the chemists who have studied the composition of the air.
17. Why was argon never before found?
18. Who were the chemists who discovered argon?
19. How was it finally discovered?
20. What are the characteristics of argon?
21. How does it differ from nitrogen?
22. How is the water vapor of the air constantly renewed?
23. On what does humidity depend?
24. What is the relation between the water vapor of the air and rain, snow, frost, or dew?

25. Where does the beading of moisture which appears on the outside of a glass of ice water in summer come from?

26. On a fair day how much water vapor may the atmosphere hold?

27. On a dog-day how much water vapor may there be in the air?

28. Why do wet clothes dry better on a clear day than on a humid day?

29. Why do not the gases of the atmosphere rest on each other in layers?

30. What is the property of diffusion?

31. What is the appearance of liquid air?

32. At what temperature does liquid air boil?

33. What happens to substances dropped into liquid air?

34. What is the definition of a mixture?

35. What is the definition of a compound?

36. How do we know from these two definitions that air is a mixture and not a compound?

37. What is the difference in the action of magnesium in air and in oxygen?

38. What is the difference of the combination of nitrogen with magnesium in air and in oxygen?

39. What is the action of oxygen in the air?

40. What is the action of nitrogen in the air?

41. What is the action of argon in the air?

42. What is the action of carbon dioxide in the air?

43. What is the action of water vapor in the air?

CHAPTER XVII

Gas Laws

1. On what laws do many of our conceptions of chemistry depend?

2. Describe the contraction and expansion of gases.

3. What every-day habit of life is affected by these gas changes?

4. What two elements are necessary for a fire?

5. How do we get air for a fire?

6. What is suction?

7. Explain the action of suction in drinking lemonade through a straw.

8. Explain how air is pulled up a chimney.

9. What is the weight of the air in a used chimney compared with the weight of the air outside?

10. State the law of Charles.

11. What per cent of itself does every true gas expand or contract for each degree Centigrade?

12. Illustrate this principle by a fleet of 273 ships.

13. If we have 273 cc. of gas at zero degree C., how many will we have at 473 degrees?
14. How many will we have at 73 degrees?
15. How many will we have at –273 degrees?
16. Why can a gas never disappear?
17. What happens to all gases before they reach –273 degrees?
18. What is absolute zero?
19. What is the value of absolute zero in Centigrade degrees?
20. What do N.T. and S.T. mean?
21. Why is a barometer a balance?
22. Explain the construction of a barometer.
23. Why does the barometer rise in fair weather and fall in foul?
24. Why must we consider pressure in calculating the weight of any gas?
25. State the law of Boyle or Mariotte?
26. Give this law as a proportion.
27. Give this law in terms of gases acted upon by air.
28. What is N.P.?
29. What is N.T.P. or S.T.P.?
30. State the law of Gay Lussac.
31. Illustrate this law.
32. State Avogadro's Hypothesis.
33. Illustrate this hypothesis by means of barns filled with swallows.
34. Illustrate this hypothesis by comparing a liter of chlorine and a liter of hydrogen.
35. What conditions must always be fulfilled to have this hypothesis correct?
36. How widely does this hypothesis apply?
37. Why are the gas laws of such importance in the study of chemical theory?

CHAPTER XVIII

SOIL

1. What were the four elements of the ancients?
2. What is water?
3. What is fire?
4. What is air?
5. What metals and non-metals are found on the crust of the earth?
6. What is the chief use of the soil?
7. How have the uses of the metals and non-metals of the earth been increased?
8. What is the soil?

9. How do all living things directly or indirectly depend on the soil?

10. What materials for their structure can vegetable organisms take from the air?

11. What other elements are most essential to the growth of plants?

12. On what does the fertility of the soil depend?

13. What makes the difference between fertile and barren soil?

14. How can a fertile soil become a barren soil?

15. Explain fertilization by manures.

16. What chemical action is going on in rotting manures?

17. Explain the various fertilizing values of manure, guano, nitrates, and lime.

18. What is the purpose of the rotation of crops?

19. What is the reason for making soil analyses?

20. Why is the fear of general starvation disappearing?

21. What necessary assistance beside providing food does soil supply to the plants?

22. What is earth?

23. Show how the character of the soil depends on the character of the rocks which are or have been present.

24. What part of the structure of the soil is most commonly found?

25. What is sand?

26. What are quartz and flint?

27. What is meant by saying "as hard as flint"?

28. How is the flintlock gun exploded?

29. What is the solubility of silica?

30. Where is silica found in plant life?

31. What compound does clay chiefly contain?

32. What other compounds does clay contain?

33. In what respect does aluminum differ from the other elements of the soil?

34. What is granite?

35. Explain the composition of quartz.

36. Explain the composition of feldspar.

37. Explain the composition of mica.

38. How is clay produced from granite?

39. What is kaolin?

40. Give some of the precious stones found in the crust of the earth.

CHAPTER XIX

SULPHUR

1. Why did the ancients believe that the center of the earth was filled with fire and brimstone?

2. Why was this especially likely to be believed by the people of Sicily and Southern Italy?

3. Where does sulphur come from beside North America and where is it found in North America?

4. What is a stratified formation?

5. Where is sulphur found in such a formation?

6. How is it found in living beds?

7. Why is an egg or mustard spoon blackened?

8. Why do silver or copper ornaments in the house turn dark?

9. Why does wheat turn black?

10. Name two vegetables that contain sulphur.

11. How much sulphur is there in the human body?

12. What are sulphides?

13. Where is lead sulphide found?

14. Where is copper sulphide found?

15. Where is mercury sulphide found?

16. Give the common name of these three sulphides.

17. What is the special use of iron pyrites?

18. What is its common name?

19. Give the properties of sulphur.

20. Tell what happens to a test tube filled with sulphur which is heated over a gas burner.

21. What chemical change takes place in heating sulphur?

22. What is an allotropic modification?

23. What are the four allotropes of sulphur?

24. Describe prismatic sulphur.

25. Describe rhombic sulphur.

26. Describe plastic sulphur.

27. Describe flowers of sulphur.

28. What is milk of sulphur?

29. How is it produced?

30. What is the test for sulphur?

31. What is the common oxide of sulphur?

32. What group of compounds does the union of metals with sulphur produce?

33. What group of compounds does the union of metals with oxygen produce?

34. What happens when heated copper or iron wire is thrust in sulphur vapor?

35. How is sulphur separated from its impurities?

36. Give the two parts of the process of manufacturing sulphur.

37. Give three commercial uses of sulphur.

38. How is sulphur used in the match?

39. How is it used in rubber manufacture?
40. How is it used as a germicide?

CHAPTER XX

SULPHUR DIOXIDE. SULPHUR TRIOXIDE. HYDROGEN SULPHIDE

1. How is the straw of the straw hat bleached?
2. What other substances are bleached by sulphur dioxide?
3. Why does an old straw hat turn to a dull yellow?
4. What is the advantage of sulphur dioxide as a bleach?
5. What is its use in paper-making?
6. Where do you commonly get the odor of sulphur dioxide?
7. What is the action of sulphur dioxide on a flame?
8. How can sulphur dioxide be used to put out a fire in a chimney?
9. What is its general action on flame?
10. What relation is there between this action and the use of burning sulphur as a disinfectant?
11. How can sulphur dioxide be prepared?
12. What is the solubility of sulphur dioxide? How can this be shown?
13. Name some other properties of sulphur dioxide.
14. How can we get a steady stream of sulphur dioxide in the laboratory?
15. Give the reaction for preparing sulphur dioxide?
16. What is the action of sulphur dioxide on the color of flowers?
17. What is the chief importance of sulphur trioxide?
18. What is one of the most striking characteristics which gaseous oxides of the non-metals possess?
19. Name several gaseous oxides which act alike when added to water.
20. What acids are formed from these oxides?
21. What is an acid anhydride?
22. What difference is indicated by the use of the endings -ous and -ic?
23. What is the composition of sulphurous acid?
24. Give the reaction for its formation?
25. What is a radical?
26. How is the radical of sulphuric acid formed?
27. What is the difference between the endings -ite and -ate?
28. What is the radical of sulphuric acid?
29. How can we make sulphuric anhydride from sulphurous anhydride?
30. What is meant by a hygroscopic body?
31. Give the various names of hydrogen sulphide?
32. Where is hydrogen sulphide found in nature?
33. What is the most important of the properties of this compound?

34. Name some metallic sulphides.

35. How can we form lead sulphide?

36. Write the reaction for the formation of copper sulphide from hydrogen sulphide?

37. How can we make hydrogen sulphide?

38. What are the properties of sulphuretted hydrogen?

39. What precaution must be taken in its making?

40. What is the action of hydrogen sulphide on white lead?

41. What is the common test for hydrogen sulphide?

CHAPTER XXI

SULPHURIC ACID AND THE SULPHATES. SULPHUROUS ACID AND THE SULPHITES

1. Where does our modern paper originate?

2. How were papers formerly made?

3. How are they made to-day?

4. What organic matters are used for paper?

5. What common property must every substance possess which is used for paper-making?

6. Of what does paper consist?

7. How is wood pulp made?

8. Why is wood pulp which has not been treated chemically not used exclusively for good paper?

9. What processes are used for preparing chemical wood pulp?

10. Which process is the most used?

11. How is the sulphite process for wood pulp carried on?

12. What does this process do to the fibers of the wood?

13. What mixture do ordinary papers contain?

14. Describe the making of paper from the pulp to the paper machine.

15. Describe the action which goes on from the time the pulp reaches the paper machine until it comes off in the final sheet.

16. On what basic property do all these changes depend?

17. How are sulphites formed?

18. Give the reaction for the forming of sulphites.

19. Give the barium chloride test for sulphurous acid.

20. Why is sulphurous acid a reducing agent?

21. What are the characteristics of sulphurous acid?

22. Give some of the uses of sulphuric acid?

23. How does the manufacture of sulphuric acid compare with that of some of the other acids?

24. What is the action of sulphuric acid on the metals?

25. Give the reaction for the effects of sulphuric acid on zinc.

26. What two methods are used for making sulphuric acid?

27. Give the three names of the contact process.

28. Give the four parts of the contact process.

29. Where do we find the beginnings of the English or chamber process?

30. How did Valentine make sulphuric acid?

31. How is the process of to-day different from the old process?

32. Where is the sulphuric acid used in both processes chiefly obtained?

33. Explain the action which sulphur dioxide undergoes, from the time it enters the lead chambers until it comes out as sulphuric acid.

34. How does the hygroscopic action of sulphuric acid enable it to be of service in drying gases?

35. Why is sulphuric acid particularly available for producing most of the other acids?

36. Why is sulphuric acid always used in electrical actions?

37. What phenomena is observed when sulphuric acid is mixed with water?

38. Why should sulphuric acid always be poured on water?

39. What is the effect of sulphuric acid on organic substances like cloth or wood?

40. Give the reaction for the formation of a sulphate.

41. Give the barium chloride test for sulphuric acid, and give the reaction for this test.

CHAPTER XXII

PHOSPHORUS AND PHOSPHATES

1. How are most of the lamps and the fires of the world lighted?

2. How did our ancestors light their fires?

3. Why is phosphorus used in the match?

4. How is phosphorus kept in its elementary form?

5. Why should it not be handled with bare hands?

6. Why does the friction match light when struck?

7. What are the two allotropes of phosphorus?

8. Describe the difference between these allotropes.

9. Why is it difficult to obtain pure phosphorus?

10. How can phosphorus oxidize under water?

11. Into what form does yellow phosphorus tend to change? Compare yellow and red phosphorus as regards odor, solubility, poisonous qualities, and phosphorescence.

12. What were the first matches.

13. Why was potassium chloride put on these matches?

14. Why is friction used in igniting a match?

15. How were the first friction matches made?
16. How are the matches of the present day constructed?
17. Why do they use sulphur as well as phosphorus?
18. What is the danger of friction matches?
19. Why are safety matches better than friction matches or common matches?
20. What are the two requirements for lighting a safety match?
21. What necessary action takes place in lighting most safety matches?
22. Why do fire departments encourage the use of safety matches?
23. From what sources is phosphorus obtained?
24. How is it made from bone-ash?
25. How is it made from phosphate?
26. How is it made from calcium phosphate?
27. Explain the manufacture of phosphorus or the way of changing phosphates into yellow phosphorus.
28. Why is phosphorus essential to life?
29. What is phosphoric acid?
30. How do phosphates come into our foods?
31. What effect do the compounds of phosphorus have upon the production of plant life?
32. Name the two acids of phosphorus.
33. What common compound of the kitchen contains phosphorus?
34. What are the essential principles of baking powders?
35. What is the action of the carbon dioxide gas evolved in rising bread?
36. What phosphate is generally used in phosphate baking powders?
37. What compounds besides carbon dioxide are formed by the action of phosphate baking powders?

CHAPTER XXIII

THE THEORY OF SOLUTIONS AND THE EFFECT OF ELECTRICITY ON CHEMICAL ACTION

1. Give the names of some of the students who investigated the relation between electrical and chemical action.
2. How is electricity generated?
3. Give two types of electrical generators.
4. How does the electric current flow?
5. How does the electric current pass into the arc light?
6. How does the current leave the battery and enter it again?
7. How does the current enter and leave a solution which conducts electricity?
8. Define an anode and a cathode.

9. Define electrodes.

10. What are electrical conductors?

11. What is likely to be the composition of such conductors?

12. What difference in action appears between a metallic conductor and a conducting solution when the current is passed through them?

13. What happens when an electric current is passed through a conducting solution?

14. What is polarity?

15. Define non-electrolytes.

16. Define electrolytes.

17. What type of bodies are commonly found to be non-electrolytes?

18. What type of bodies are commonly electrolytes?

19. How can we distinguish experimentally between electrolytes and non-electrolytes?

20. Who brought forward the Theory of Electrolytic Dissociation?

21. What is the difference between the condition of an electrolyte like sodium chloride in solution and in the solid state?

22. What is dissociation?

23. What occurs when molecules dissociate?

24. What is an ion?

25. What should we see if we could observe salt in solution in a glass of water?

26. Explain the interchange between the dissociated and the undissociated part of such a solution of salt.

27. Explain the production of ions according to Talbot and Blanchard.

28. What is the relation between the positive charges in the solution and the negative charges?

29. What is the relation between the ionic theory and the combination of elements with regard to the theory of valence?

30. Explain the formation of silver chloride, copper chloride, and ferric chloride according to this theory.

31. Explain two (*acid plus base equals salt plus water*) reactions in forms of the ionic theory.

32. What is the difference between writing molecular and ionic symbols?

33. Why is not the symbol of water in reactions given in text written with plus and minus signs?

34. Give the former definition for valence.

35. Define a monovalent, bivalent, and trivalent ion.

36. Does a relation possibly exist between the valence of atoms and compounds in the ionic and the molecular states?

37. How does a single fluid galvanic cell generate electricity?

38. Explain the construction of a single galvanic cell.

39. Explain the decomposition of sulphuric acid in a single fluid galvanic cell.

40. Define an anion. Define a cation.

41. What actions occur when the ions in a solution reach electrodes?

42. Explain electrolytic action in silver plating.

43. Give Faraday's law:

44. Explain it by illustration of three conducting solutions in a constant circuit.

CHAPTER XXIV

CARBON

1. Where do diamonds come from?
2. Where is graphite found?
3. Where is charcoal prepared?
4. How is lampblack commonly made?
5. Where is gas carbon and coke manufactured?
6. How is coal obtained?
7. Give three allotropes of carbon.
8. How do we know that the diamond is pure carbon?
9. How does graphite differ from the diamond in composition?
10. How are diamonds made in nature?
11. How are they made artificially?
12. What chemists prepared artificial diamonds?
13. What is the heat of the electric furnace?
14. What two elements did Moissan use in preparing diamonds?
15. What action went on in the crucible?
16. Why was this process not a commercial success?
17. Give the properties of the diamond.
18. For what is graphite chiefly used?
19. What other commercial uses does graphite possess?
20. How are lead pencils made?
21. What are the properties of graphite?
22. What forms of carbon are grouped under the heading of amorphous carbon?
23. Give two uses for lampblack.
24. How is wood charcoal prepared?
25. What is destructive distillation?
26. Why is wood charcoal used to give a hot fire?
27. Why are wood posts charred before they are set in the ground?
28. What is the use of charcoal as a deodorizer and filter?
29. What is the danger of the common household filter?
30. How is charcoal used in sugar refining?

31. Give the different characteristics of coal.

32. What was the original state of the coal beds such as are found in Pennsylvania and Wales?

33. Explain the gradual formation of coal.

34. What is the radical of the cyanides?

35. What dangerous characteristic belongs particularly to this group?

36. What is the acid of the group?

37. Where is prussic acid found in nature?

38. What is the most important salt of hydrocyanic acid?

39. Where is it used?

40. Why is it a dangerous reagent?

41. Give some of the early uses of the cyanides.

42. What are the fulminates?

CHAPTER XXV

ORGANIC CARBON

1. In what respect does carbon differ from other chemical elements?

2. Name some of the natural products in which carbon is found.

3. Define organic chemistry.

4. Define inorganic chemistry.

5. What fundamental difference is there between the two divisions?

6. What laws do both the divisions obey?

7. What was the old conception of organic chemistry?

8. What was the old definition of organic chemistry?

9. How did Wöhler destroy the old theory?

10. Why did the old division continue after Wöhler's discovery?

11. Give a broad definition for organic chemistry.

12. What is the reason for the separation of the two divisions of chemical knowledge?

13. Compare organic and inorganic compounds with regard to their action with electricity, with the elements which compose them and their general properties.

14. Name six inorganic compounds.

15. Name six organic compounds.

16. On what does the difference of inorganic and organic chemistry chiefly depend?

17. How is methane formed in nature?

18. What is the action which goes on when marsh gas is produced in a stagnant pool?

19. What is the chemical composition of methane?

20. What is valence?

21. What is a monovalent substance?

22. What is a bivalent substance?
23. What is a trivalent substance?
24. What is the valence of carbon?
25. How is methane made in the laboratory?
26. Give the characteristic properties of methane.
27. How is wood alcohol commonly produced?
28. What are the properties of wood alcohol?
29. What is its action as a solvent?
30. What is the composition of methyl alcohol?
31. What can you say of the basic action of methyl alcohol?
32. What is an aldehyde radical?
33. What is the chief use of formaldehyde?
34. What is the danger of this use of formaldehyde?
35. How is formaldehyde prepared?
36. What is formalin?
37. How is it tested?
38. What is the radical of the organic acids? How is this radical formed?
39. Where is formic acid found in nature?
40. How was it formerly prepared? How is it prepared in the laboratory?
41. How is methyl ether prepared?
42. What are chloroform and iodoform?

CHAPTER XXVI

ETHANE

1. Where is ethane found in nature?
2. Give the characteristics of ethane.
3. Give some of its compounds.
4. What is the chemical composition of ethane?
5. How are the hydrocarbons of this series prepared?
6. What is the chemical composition of propane?
7. What is the highest compound of this series which has been produced up to the present time in the laboratory?
8. How many compounds of this type can be formed?
9. What is the chemical composition of spirits of wine (ethyl alcohol)?
10. How is ethyl alcohol produced?
11. How is it separated from impure alcohol?
12. Describe fractional distillation.
13. Describe the separation of alcohol from a mixture of water and alcohol.
14. What are the boiling and freezing points of alcohol?

15. What uses are made of the low freezing-point of alcohol?
16. What occurs when alcohol burns?
17. What is the test for alcohol?
18. What is an anæsthetic?
19. What was the first anæsthetic produced?
20. Give the names of the discoverers who brought about the use of sulphuric ether.
21. Why are anæsthetics of such value?
22. What is the chemical constitution of sulphuric ether?
23. Describe the esters and their properties.
24. What substances are obtained from the destructive distillation of wood?
25. What are the properties of pyroligneous acid and what three compounds does it contain?
26. What are the acetones and what radical do they contain?
27. How is sulphuric ether made?
28. What can you say of the insolubility of sulphuric ether?
29. How is acetic acid prepared?
30. What are the uses of acetic acid?
31. How are acetates prepared?
32. Give some of the uses of the acetates.

CHAPTER XXVII

Fermentation

1. What are the three most important divisions of the plants of the air?
2. Where are these plants found?
3. How can we see these plants?
4. Where do we find the molds?
5. What is the commonest form of yeast?
6. What is a yeast cake?
7. What common food shows bacterial life?
8. What is the size of bacteria?
9. How do bacteria act when they meet waste organic matter?
10. How do modern sewage plants take advantage of this action of the bacteria?
11. What is the action of the bacteria on the woody fiber of the trees?
12. To what is decay due?
13. Why is decay a necessity of natural order?
14. Name some agricultural processes which are aided by bacterial decay?
15. Give two of the general names for bacteria, yeasts and molds.

16. In what general botanical class do they belong?

17. How do they differ among themselves?

18. Why do micro-organisms flourish in the foods of man?

19. On what principle does the large part of the science of food preservation depend?

20. How is this shown in jelly-making and in the preservation of fish and meats?

21. Why is salt food impossible as a steady diet?

22. How is the change from soft to hard cider brought about?

23. Describe the action of the yeast in the cider press.

24. What two compounds always appear in the presence of fermentation?

25. How are alcoholic drinks obtained?

26. Describe the yeast plant and the yeast cake.

27. How do yeasts grow?

28. Describe the raising of bread.

29. What chemical compounds are produced when bread rises?

30. What effect does the yeast have upon the flavor of bread?

31. How is hard cider changed into cider vinegar?

32. What is the structure of the bacteria?

33. How do they grow?

34. Why does hard cider afford a fertile soil for bacteria?

35. Describe the process of changing hard cider to vinegar.

36. How does the oxidation of the derivatives of the ethane group resemble the oxidation of the derivatives of the methane group?

37. What disease germs are found among bacteria?

38. How can we guard ourselves against these bacteria?

39. Why is it especially necessary to guard against bacterial life in the milk supply?

CHAPTER XXVIII

Foods

1. What energy controls the natural resources of the earth?

2. How does man obtain his energy?

3. What are foods chemically speaking?

4. What elements are found in the body structure?

5. What elements then must the food supply?

6. What two requirements have food beside the supplying of the body structure?

7. What are the five classes of nutrients?

8. Why do we consider foods in terms of these classes instead of terms of bread and cheese for example?

9. What three factors must always be kept in mind with regard to foods?
10. What is the action of carbon and hydrogen in the body?
11. Where are the carbohydrates found?
12. What is the chief value of the carbohydrates?
13. Where are the fats found?
14. What is the chief value of the fats?
15. What may happen to the excess of fats and carbohydrates taken in the diet?
16. How is bone structure provided for?
17. What are the two necessary services of water?
18. What are body cells?
19. How do the body cells work?
20. What nutrient supplies the structure of the cells?
21. What happens to the excess of proteid beyond that required for building up the body tissue?
22. Why is proteid an absolute necessity for the diet?
23. What is the chemical composition of proteid?
24. Give the percentage of the nutrients in bread and milk.
25. Give the percentage of the nutrients present in peas and in meat.
26. How was soap formerly made?
27. What differences are there between the method used then and those which are in use to-day?
28. What three organic substances are chiefly found in soap fats?
29. What happens first when you put fat in a kettle with an alkali and boil the mass?
30. What is the second action?
31. What compounds form when sodium hydroxide is used?
32. What compounds form when potassium hydroxide is used?
33. How are soaps brought out of solution?
34. How can we make soft and castile soap?
35. On what does the cleansing action of soap depend?

CHAPTER XXIX

CLOTHES

1. What are the threads of cloth?
2. What are silk, wool, cotton, and flax?
3. What are vegetable fibers?
4. What elements appear in them?
5. What is the composition of animal fibers?
6. How do animal fibers appear under the microscope?

7. What commercial use exists for cotton and flax beside the textile use of fiber?
8. For what is cottonseed oil used?
9. For what is linseed oil used?
10. How do the drying oils act?
11. How are these oils separated?
12. What is the appearance of the press cake?
13. For what is it used?
14. Give the divisions of vegetable fibers.
15. How do cotton fibers grow?
16. How do flax fibers appear under the microscope?
17. How is flax fiber prepared?
18. How does the silkworm produce silk?
19. How is the silk, taken from the silkworm, made ready for manufacture?
20. What class of threads come under the general class of wool?
21. What animal gives the greatest part of the wool produced?
22. Describe the appearance of fibers of wool.
23. What are worsted goods?
24. What are woolen goods?
25. What is the chemical composition of wool?
26. How is wool prepared for manufacture?
27. How is cotton prepared for manufacture?
28. How is linen bleached?
29. How are animal fibers bleached?
30. Who discovered the first aniline dye?
31. What are aniline dyes made from?
32. What are the essential principles of dyeing?
33. What is the difference in appearance between animal and vegetable fibers?
34. Describe the dyeing of white silk with picric acid.
35. Describe the dyeing of vegetable fibers with picric acid.
36. What is the reason for the different action of dyes on animal and vegetable fibers?
37. Explain the action of dyeing by means of a mordant.
38. Define a mordant and a lake.
39. Describe the action of alizarin upon cloth.
40. Describe the action of mordanting.
41. Where is alizarin found in nature?
42. What was the effect of synthetic alizarin on the production of madder?
43. Where is natural indigo found?
44. What was the effect of the production of synthetic indigo on the indigo plantation?

CHAPTER XXX

HEAT AND LIGHT IN CHEMISTRY

1. Give two household experiments in heat.

2. What two classes of heat phenomena accompany chemical action?

3. How are chemical reactions measured?

4. What is a calory?

5. How many calories should we use in raising ten grams of water ten degrees? What substance requires the most heat to raise it through one degree?

6. What is specific heat?

7. If the specific heat of a substance is .4, how high will a body rise in temperature when ten calories are used in heating it?

8. How much heat is evolved in forming eighteen grams of water?

9. How much heat is absorbed in breaking down eighteen grams of water?

10. Give the relation of this heat equation to the law of conservation of energy.

11. State the law formulated.

12. What is practically the most effective exothermic reaction?

13. What three substances distill over when coal is destructively distilled?

14. What is the use of coal gas? What is the use of ammoniacal liquor?

15. What is the composition of ordinary coal gas?

16. Which gas in coal gas furnishes light, and which furnishes heat?

17. Describe the production of gas in a gas plant.

18. Why do corporations test the fuel they receive?

19. What is the ignition-point?

20. Is this point fixed for all substances?

21. What effect does the ignition-point have upon exothermic and endothermic action?

22. When does flame occur in combustion?

23. What is the construction of the hydrogen flame burning in oxygen?

24. What is the reason for the cone shape of the flame when it issues from a round tube?

25. What forms the inner cone under such circumstances?

26. What compounds form when methane and ethane burn?

27. What three portions appear in a coal gas flame? How does admitting air to the base of the flame change the construction of the flame?

28. What theories are there concerning the reason for the light given out by the flame?

29. Why does the flame sometimes strike back?

30. What is the oxidizing flame of the Bunsen burner? What is the reducing flame of the Bunsen burner?

31. How can the action of the oxidizing and reducing flames be shown experimentally?

32. What piece of apparatus constructed by man gives the highest heat?

33. Who discovered the electric furnace?

34. Describe the construction of the electric furnace.

35. Describe the passing of the electric current through the electric furnace.

36. What relation has heat measurement to food?

37. Give three common examples of the effect of light upon chemical compounds.

38. What type of light change do we understand especially well?

39. On what substances does light act with especial force?

40. What action takes place when light strikes a silver salt on a photographic plate?

41. What action takes place when the photographic plate is developed?

42. What is the difference between a negative and a positive?

43. How are prints made from negatives?

CHAPTER XXXI

IRON AND COPPER

1. What is the esential principle of the reduction of iron ore to metallic iron?

2. Give the five properties which especially make iron useful to man.

3. Give an illustration of each of these processes.

4. What is the special advantage of wrought iron?

5. Give the commercial use of wrought iron.

6. How does wrought iron differ from steel and cast iron in composition?

7. How does steel differ from wrought iron and cast iron in its composition?

8. What would the name manganese steel imply?

9. What did the story of "The Talisman" illustrate with regard to the possibilities of steel?

10. What is the tempering of steel?

11. Why did the sword of Saladin act so differently from Richard's sword?

12. How can we get brittle and hard steel?

13. How can we get soft and tough steel?

14. Give the different types of steel which we obtain at different temperatures.

15. What is the effect of sudden cooling on other types of iron besides steel?

16. What is cast iron?

17. How does it differ from wrought iron and steel?

18. How is it used commercially?

19. What is the simplest method of reducing soft, pure, iron ores?

20. Where is iron found?

21. Name the different ores of iron.

22. Which ores are chiefly used for producing metallic iron?

23. What peculiar power does magnetite possess?

24. Where is iron found in the United States?

25. What are the three necessities for iron-making?

26. Why is Birmingham, Ala., a large producer of iron?

27. How is iron smelted in the Pyrenees?

28. Why do we require lime for iron-making?

29. What is the action of the lime on the ore?

30. What are the three main features of iron smelting?

31. Describe a blast furnace.

32. Describe the process used on iron ore in a furnace. What are tuyeres and how are they used?

33. How is a blast furnace kept hot?

34. What mixture is thrown into a blast furnace?

35. Explain the reducing action of carbon and the change through which it passes in reducing iron.

36. Explain the chemical action of limestone in the blast furnace.

37. How are iron slags taken off from a furnace?

38. What two groups of salts does iron form?

39. What is the test for iron?

COPPER

40. What is the chief use of copper?

41. Why is copper used for electrical work?

42. Describe the properties of copper.

43. Give a test for copper.

44. How is copper used commercially?

45. What is an alloy?

46. What is the composition of brass, bronze and coin nickel?

47. For what is copper sulphate used?

48. Where is copper found?

49. How is it reduced from its ores?

CHAPTER XXXII

Sodium. Potassium. Calcium. Aluminum. Gold. Silver

1. What is the flame test?
2. What is the flame test for sodium?
3. How was metallic sodium first prepared?
4. How is it prepared to-day?
5. What is the action of sodium on water?
6. How is yellow fire made?
7. What is borax?
8. What is boron?
9. Where is borax found?
10. What is its use in commerce?
11. What use of borax may prove dangerous to the community?
12. How is metallic potassium made?
13. Where are potassium salts found in nature?
14. What are the properties of potassium?
15. How is sodium and potassium kept for use?
16. What is the action of potassium on water?
17. What is lime?
18. What is the flame test for potassium?
19. What is the use of calcium oxide in commerce?
20. What is the composition of the earth's crust?
21. What are the properties of metallic calcium?
22. How is the calcium flame produced?
23. How is metallic calcium prepared?
24. How is aluminum prepared?
25. What has been the change in the cost of aluminum and how was it brought about?
26. On what properties does the value of aluminum depend and how is it used in commerce?
27. What jewels are compounds of aluminum?
28. What is corundum and for what is it used?
29. What are the two great uses of gold and silver?
30. Where is gold found?
31. How is gold extracted from quartz ores?
32. Give the properties of gold.
33. What are the alloys of gold?
34. What is the effect of alloying gold?
35. Describe the measure of gold in carats.
36. What is the great advantage of gold and silver for use in coins and jewelry?
37. How is gold extracted from its ores?

38. What is an amalgam?
39. Give the properties of silver.
40. Give the test for silver.

CHAPTER XXXIII

PAINTS. LEAD. ZINC. TIN. ARSENIC. ANTIMONY

1. What are paints?
2. Why are paints so universally used?
3. Where do pigments come from?
4. What is the difference between pigments and dyestuffs?
5. How are pigments applied?
6. What is the essential action of pigments in regard to light?
7. What is white lead and what is covering power?
8. What are the four great advantages which white lead possesses?
9. How was white lead formerly prepared?
10. How is white lead now prepared and how much white lead can be made at a time?
11. What are the two great disadvantages of white lead?
12. How is zinc oxide made? Compare zinc oxide and white lead.
13. How are blue, red, and yellow pigments prepared?
14. What are the two chief advantages of lead? How is lead used commercially?
15. Give the properties of lead?
16. How can lead be reduced from its oxides?
17. How can lead be reduced from its ores?
18. What is the danger of lead pipe in water systems?
19. Why should water be allowed to flow from water pipes before drinking?
20. In what ways is zinc more useful than lead?
21. What is the comparative density of zinc and lead?
22. What are some of the commercial uses of zinc?
23. How is zinc prepared from its ores?
24. What are the properties of zinc?
25. How are tin pans made?
26. Why do they cover iron pans with tin?
27. What is solder? What is pewter?
28. Give the properties of tin. How is it reduced from its ores?
29. What is a metalloid?
30. In what way is arsenic like the metals? In what way is it like the non-metals?
31. What are the oxides of arsenic?
32. What is arsine?

33. Why are arsenic wall papers or hangings dangerous?
34. How is arsenic found in nature and how is it prepared?
35. What is the chief value of antimony?
36. How is type-metal made?
37. Give the properties of antimony.
38. Compare stibine and arsine.
39. What oxide is formed when antimony burns?
40. How is antimony reduced from its ores?

INDEX

Note Carefully. Throughout this index where any heading refers to two or more different pages, that page on which the given subject receives its fullest treatment is given first and is printed in blacker type.

For example, the heading "Hydrochloric acid. **91**, 52, 89, 126." On page **91** and the immediately succeeding pages will be found the general discussion of hydrochloric acid, its occurrence, properties and preparation. The other page numbers given above refer to hydrochloric acid in its relation to other compounds and elements.

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Yellow prussiate of potash. 259.

Z

Zinc, 359.
Zinc chloride, 101.
Zinc white, 355.

WELLCOME LIBRARY INSTITUTE

ATOMIC WEIGHTS AND SYMBOLS OF IMPORTANT ELEMENTS

NAME	SYMBOL	ATOMIC WEIGHT	
		Accurate O = 16	Approximate (to be used for all calculations)
Aluminum	Al.	27.1	27
Antimony	Sb.	120.2	120
Argon	A.	39.9	
Arsenic	As.	75.0	75
Barium	Ba.	137.37	137
Bismuth	Bi.	208.0	208
Bromine	Br.	79.92	80
Boron	B.	11.0	11
Cadmium	Cd.	112.40	112
Calcium	Ca.	40.09	40
Carbon	C.	12.00	12
Chlorine	Cl.	35.46	35.5
Chromium	Cr.	52.1	52
Cobalt	Co.	58.97	59
Copper	Cu.	63.57	63.5
Fluorine	F.	19.0	19
Gold	Au.	197.2	197
Hydrogen	H.	1.008	1
Iodine	I.	126.92	127
Iron	Fe.	55.85	56
Lead	Pb.	207.10	207
Lithium	Li.	7.00	7
Magnesium	Mg.	24.32	24
Manganese	Mn.	54.93	55
Mercury	Hg.	200.0	200
Nickel	Ni.	58.68	58.5
Nitrogen	N.	14.01	14
Oxygen	O.	16.00	16
Phosphorus	P.	31.0	31
Platinum	Pt.	195.0	195
Potassium	K.	39.10	39
Silicon	Si.	28.3	28
Silver	Ag.	107.88	108
Sodium	Na.	23.00	23
Strontium	Sr.	87.62	87.5
Sulphur	S.	32.07	32
Tellurium	Te.	127.5	
Thorium	Th.	232.42	
Tin	Sn.	119.0	119
Tungsten	W.	184.0	
Zinc	Zn.	65.37	65.4